A SPECIAL MESSAGE TO PARENTS

When your son, or daughter, asks for a CHEMCRAFT Outfit you can be sure that he is asking you for more than a mere toy. He is asking you to make an important and worthwhile investment in his future.

Many men and women who received CHEMCRAFT Outfits years ago have applied their knowledge of chemistry to the vocations and businesses they now pursue. For these sets are an interesting introduction to this wonderful field of science, stimulating interest in further scientific studies, and giving an excellent background for courses taken in elementary schools, high schools, and colleges. Our correspondence contains many laudatory letters from parents, teachers, and school principals. We are sure that you will find that your son or daughter will not only receive hours of wholesome fun from these sets, but they will also be acquiring basic scientific knowledge that will prove quite useful to them in later years.

We want you to know that CHEMCRAFT Outfits are completely safe. They contain no poisons or explosives, nor will any combination of the chemicals result in an unsafe or unpleasant situation. These sets, and their contents and experiments, must adhere to rigid City, State, and Federal laws.

CHEMCRAFT Outfits and SCIENCECRAFT MICROSCOPES are genuinely educational, while giving boys and girls many hours of constructive enjoyment and fun.

Boys and Girls--Send for your FREE Copy of the CHEMCRAFT NEWS

National Headquarters of the Chemcraft Science Club invites you to send for your copy of the new CHEMCRAFT NEWS. Clip the coupon below, and mail it today. This latest issue of the CHEMCRAFT NEWS is chock full of interesting articles and letters—many of which have been sent in to us by your fellow members. Mail your coupon today!

CLIP COUPON Mail Today

Please send the latest issue of the CHEMCRAFT NEWS to me as soon as possible. I am a member of the Chemcraft Science Club and my name and address is:

Name _____ My Age is: _____
(Please print)

Street _____

City _____ State _____

REG. U.S. PAT. OFF.

THE CHEMCRAFT SCIENCE CLUB

HAND BOOK AND CATALOG

1951

The PORTER CHEMICAL COMPANY

HAGERSTOWN, MARYLAND

An Introduction to the NEW AGE OF CHEMISTRY

BOYS—here is your new CHEMCRAFT Catalog and Handbook. In it you will find the finest assortment of chemical laboratory equipment, chemical refills for your CHEMCRAFT Outfits, apparatus, books and manuals, complete CHEMCRAFT Outfits, and SCIENCECRAFT MICROSCOPES, that we have ever offered. Here you will read about the unusual Atomic Energy feature—the greatest development in the history of Chemistry Sets.

The Senior, Master, and Master DeLuxe CHEMCRAFT Outfits contain special experiments and materials, giving you the opportunity to study the tremendous new force of ATOMIC ENERGY! In an exciting, yet completely safe way you will actually have Uranium Ore with which to experiment—plus a phenominal new instrument called a Spinthariscope, atomic energy equipment, and an extremely interesting manual. No other sets offer this development—the most astounding in the history of all chemistry sets.

In addition to Atomic Energy, the sets shown here contain the many outstanding features that have made CHEMCRAFT and SCIENCECRAFT MICROSCOPES the leaders in their fields. No other sets have the quantity and quality of chemicals and apparatus, the complete manuals, and the large number of exciting experiments you will find in these Outfits. We are proud to present this entirely new and vastly improved line of sets.

CHEMCRAFT is considered by millions of boys and their parents as the world's finest introduction to chemistry. Today we stand on the threshold of a great new age in this marvelous field of science. While you explore the wonders of chemistry with CHEMCRAFT you will have fun, thrills, and adventure—and will be preparing yourself for any one of the multitude of fields where chemistry is of value in later life. Remember, you get more, you can do more, and you have more fun with CHEMCRAFT.

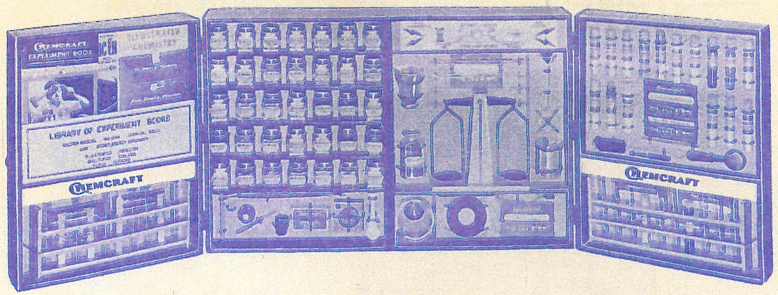
Exclusive CHEMCRAFT Features

- Atomic Energy Manual
- Atomic Energy Experiments
- Atomic Energy Spinthariscope
- Uranium Ore
- Radioactive material
- "The Story of Atomic Energy"
- Bryan Valence Blocks
- "The Story of Chemistry"
- Chemcraft Industrial Chemistry Charts
- The CHEMCRAFT Balance and Weights
- Qualitative Analysis
- Glass Blowing
- Alcohol Lamp Blow Torch
- Laboratory Construction Manual
- Chemical Magic
- Heating Stand and Screen
- Chemicals with Formulae and Names
- Gas Generating Apparatus
- Interesting and Complete Manuals
- Non-corrosive Cabinets

Not all features are found in all sets, but most, and many more, appear in the majority of the Outfits. All CHEMCRAFT sets are completely safe and harmless.

CHEMCRAFT REG. U.S. PAT. OFF.

LARGEST AND FINEST CHEMISTRY OUTFITS MADE
CHEMCRAFT ADVANCED LABORATORIES



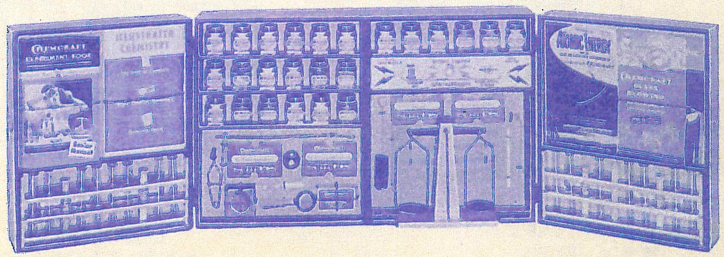
CHEMCRAFT MASTER DELUXE LABORATORY — **Eastern Price $27.50** — **Western Price $28.50**

Do you want the best? This completely new and re-designed Chemcraft Outfit is the finest, largest and most complete experimental chemistry outfit available. It represents the fulfillment of every boy chemist's dream, and is designed to provide a superb collection of chemical reagents, apparatus and equipment for home study and experimental research. The Atomic Energy feature is tops in its field—exciting and wonderful, and of real educational value. Important new features in the Master Deluxe Laboratory include the new metal Select-O-Shelves with new glass chemical containers, and the following:

"The Story of Atomic Energy"
The Chemcraft Spinthariscope
Uranium Ore
Radioactive Screen
Atomic Energy Experiments
New Chemcraft Balance and Weights
The Bryan Chemical Illustrators
The Alcohol Lamp Blow Torch
Gas Generating Apparatus
Mortar and Pestle
Glass Beaker
Chemical Graduate
Tripod and Wire Gauze
Separate Test Tube Racks
Library of Experiment Books

"The Story of Atomic Energy" is the first popular presentation of this subject ever developed specifically for boys and girls. It is extremely interesting, educational and enlightening, and gives a splendid background of knowledge on this astounding subject. The Chemcraft Spinthariscope is an amazing instrument that actually shows the emanation of atomic rays. The Uranium Ore, the Radioactive Screen, the special Atomic Energy apparatus, and other units of this Atomic Energy feature represent the most significant development ever offered in a chemistry outfit.

The Master Manual is the most comprehensive of its kind ever published, and describes and explains 857 chemical experiments with hundreds of illustrations. In addition, a complete Manual on Qualitative Analysis is included, giving 111 tests used in the analysis of various substances. The manual on Glass Blowing and the Construction of Laboratory Apparatus gives 46 experiments and 17 projects. The various instruction books describe and explain a total of 1027 scientific experiments, and all of the chemicals and equipment required for performing them are included in this outfit.



CHEMCRAFT MASTER LABORATORY — **Eastern Price $17.50** — **Western Price $17.95**

Here is the big set you have seen advertized in BOY'S LIFE, OPEN ROAD FOR BOYS, POPULAR SCIENCE, MECHANIX ILLUSTRATED, SCHOLASTIC ROTO, and many other leading magazines.

In addition to the Atomic Energy feature, this attractively arranged set contains over 80 pieces of chemicals and chemical equipment. An interesting illustrated Manual of over 100 pages, plus sections on Glass Blowing and Chemical Magic, present over 700 exciting experiments covering a wide variety of the important branches of chemistry.

The Master Laboratory outfit comes in an attractive laboratory-style wood cabinet with the new Select-O-Shelves, and contains the new Chemcraft Balance.

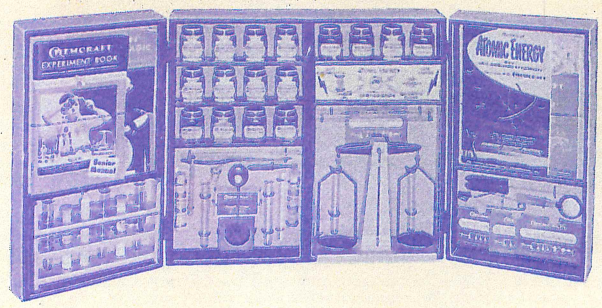
Western Prices apply to the 11 Western States

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

CHEMCRAFT REG US PAT OFF

SELECT ONE OF THESE COMPLETE OUTFITS

CHEMCRAFT HOME LABORATORIES



CHEMCRAFT SENIOR LABORATORY

Eastern Price $12.50
Western Price $12.95

Packed with a splendid assortment of chemicals, laboratory equipment and fascinating experiments, this big Senior Laboratory outfit is the leader in the Chemcraft line. It includes such outstanding features as Atomic Energy, actual specimens of Uranium Ore, "The Story of Atomic Energy," the amazing Chemcraft Spinthariscope, the Chemcraft Alcohol Lamp, Blow Torch for Glass Blowing, and a big cabinet full of other chemical apparatus, equipment and chemicals. The complete Chemcraft Manual which accompanies this outfit explains hundreds of scientific and entertaining experiments, and the separate book of mysterious experiments in Chemical Magic gives over 95 feats of magic. In the manual on Glass Blowing and The Construction of Laboratory Apparatus, the forming and blowing of glass is explained, and plans are given for making many useful pieces of laboratory apparatus. Altogether more than 600 experiments are given in the manuals accompanying this outfit. Size open 31½ x 14 inches.

CHEMCRAFT SENIOR OUTFIT

Eastern Price $7.50 Western Price $7.95

The new Senior Outfit contains 43 pieces of chemicals, chemical apparatus and equipment, instructions for Glass Blowing, and the large Senior Manual. With this fine set you can perform 471 interesting experiments, including many mystifying tricks in chemical magic. The new and improved cabinet contains a neat arrangement of all parts, and a separate laboratory-style Test Tube Rack, and a wide assortment of chemicals. Size open 21 x 14 inches.

CHEMCRAFT JUNIOR DELUXE OUTFIT

Eastern Price $4.95 Western Price $5.25

The Junior Deluxe Outfit has an attractive stand-up laboratory-type wood cabinet. More than 300 exciting experiments are described in the Chemistry Manual, and a section on Chemical Magic which tells how to perform over 70 mysterious tricks. About thirty different chemicals and pieces of apparatus each have an individual place on the shelves and in the racks with which the cabinet is fitted. Size open 20 x 10.

Western Prices apply to the 11 Western States

CHEMCRAFT REG US PAT OFF

For a Fine Start in Chemistry
CHEMCRAFT JUNIOR OUTFITS

CHEMCRAFT JUNIOR OUTFIT

Eastern Price $2.98 Western Price $3.25

Boys get a thrill when they put on a Chemcraft Magic Show. The Chemcarft Magic instructions included in this Junior Outfit tells how to perform mysterious exhibitions of chemical magic. Besides, there are 170 entertaining and educational experiments described in the regular Junior Manual. This colorful and attractive outfit contains 27 chemicals and pieces of apparatus including a laboratory-style test tube rack, test tube holder, a supply of test tubes, and other equipment. Each experiment can be repeated many times with the materials supplied with this outfit.

CHEMCRAFT BEGINNER'S OUTFIT

Eastern Price $1.98 Western Price $2.25

This dandy beginner's outfit contains 18 chemicals and pieces of apparatus, including a laboratory-type wooden test tube rack, test tubes, steel test tube holder, measures, all necessary chemicals, and a Chemcraft Instruction Manual containing clearly explained instructions for performing 151 amazing experiments. The liberal quantities of chemicals make it possible to repeat these experiments many times.

HOW TO ORDER FROM THIS CATALOG

Address all letters, checks and money orders to The Porter Chemical Company, Hagerstown, Maryland, and mail all orders to that address.

In ordering the various items you desire, be sure to give our catalog No. wherever shown, full description of the item just as it is printed in the catalog, together with size and quantity—and the proper amount of payment.

Send checks, money orders, or cash with your order. Please include 10c for postage on orders of 25c or less, include 20c for postage on orders of $1.00 or less, and include 25c for postage on orders for $2.00 or less. Send 40c for postage on all orders between $2.00 and $5.00. We pay the postage within the United States on all orders for $5.00 or over.

If you send cash please wrap carefully. United States Stamps will be accepted for amounts up to $1.00 but they must be in usable condition. Canadian or foreign stamps or coins can not be accepted. Prices and contents subject to change in conditions beyond our control.

All orders are checked before shipping and claims for shortages must be entered within 5 days from time goods are received. Be sure to look carefully through packing material when unpacking shipment, as many times articles are thrown out with the packing by the customer. We double check all packages, and wrap extremely carefully to assure you of the safe arrival of your order. All shipments are insured, but you must advise us of claims for breakage and shortages within 5 days, or we can not be responsible.

Please note: All of our chemicals and apparatus are made to rigid specifications, and are in conformance to all laws regulating their manufacture and usage. In virtually all cases our chemicals are absolutely harmless—in the case of the one or two exceptions we have been careful to mark these chemicals "poison", though this is merely an additional precaution, as the strength of the solution is usually weaker than that obtainable locally. Still as an additional precaution we request that the order for any chemical marked "poison" be signed by the parent or guardian.

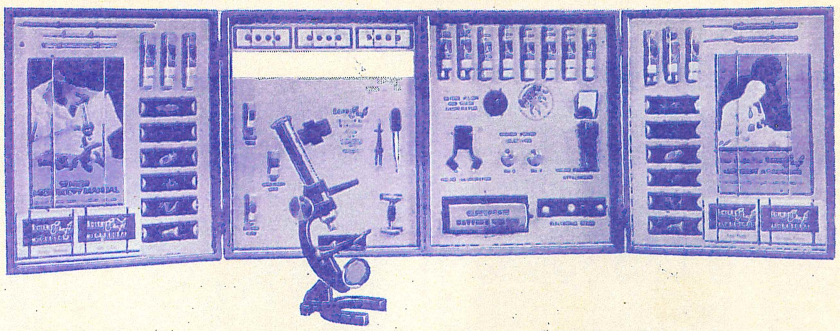
(If Resident Of State Of Maryland Include 2% For Sales Tax)

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

CHEMCRAFT

Worlds Finest Microscope Outfits for Boys

Now -- With Real Uranium Ore



SCIENCECRAFT MASTER DELUXE MICROSCOPE OUTFIT

This deluxe microscope outfit is the finest you have ever seen. Included in it you will find a complete set of attachments and accessories (see sketches below) that will aid you in performing experiments and observations that have heretofore been impossible. The large, all metal, precision built microscope has an internal arrangement of lenses, and a velvet-touch rack and pinion mechanism for focusing that are the finest we have ever offered. With the unusual features of this extra equipment you also receive specimens of real Uranium Ore, Gold Ore, and Silver Ore!

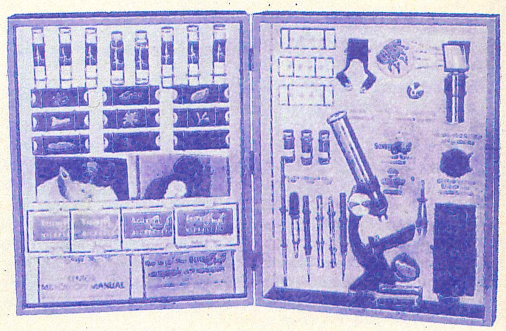
Included in this outfit is the following complete set of auxiliary attachments and accessories: 17 vials containing chemicals and specimens for observation, 7 assorted microscopy instruments including a rubber bulb pipette, a supply of lens paper, and specimen covers, two additional objective lenses of medium and high power, slides, overhead micro-illuminator for use on solid subjects, a magnifying sub-stage illuminator for use on transparent subjects, a clear view microscope projector attachment for casting large images on wall or screen, an electronic stage for observing chemical reactions, and an electronic battery holder.

A new and special feature is a series of 4 in 1 slides, with a total of 12 exciting specimens already mounted on glass slides, and ready for immediate observation.

Two highly interesting and entertaining manuals, containing full instructions for the use of the microscope and its accessories as well as many interesting observations and experiments, are included.

The red enameled Duo-form wood cabinet is 46 x 14 inches when open.

Eastern Price $21.00 **Western Price $22.00**



SCIENCECRAFT MASTER MICROSCOPE OUTFIT

Here is an excellent and very complete microscope outfit which includes a 10 inch all metal microscope, one additional higher power objective lens unit, all necessary equipment such as vials containing specimens and chemicals, glass slides, assorted dissecting instruments, a supply of lens papers and cover glasses and such outstanding accessories as the microscope projector attachment, the substage illuminator and electronic battery holder. Specimens of rare Uranium Ore, Gold Ore, and Silver Ore are also included. The large microscope manuals give complete instructions. This outfit is packed in a laboratory style wood cabinet size open 26 x 16 inches, Brilliantly enameled.

Eastern Price $15.95 **Western Price $16.50**

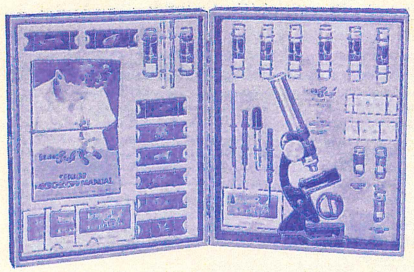
Western Prices apply to the 11 Western States

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

CHEMCRAFT

Discover the Exciting Mysteries of Nature

COMPLETE MICROSCOPE OUTFITS



SCIENCECRAFT SENIOR MICROSCOPE OUTFIT

Price $11.95 **Western Price $12.50**

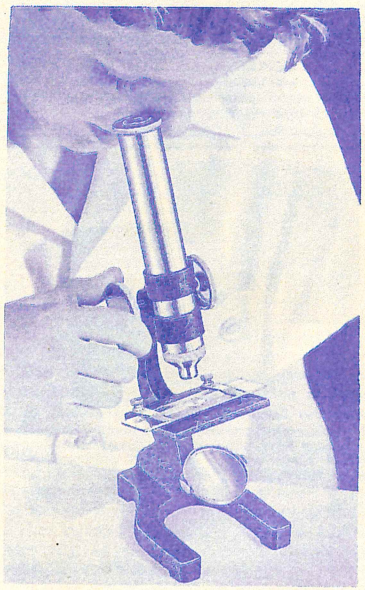
This Set offers a precision built Microscope, and a fine assortment of specimens, glass slides, and instruments — all packed in a wood laboratory-style cabinet which is 23 x 14 inches when open. An excellent Manual is also included.

This Set is an outstanding value, and a perfect addition to any boy's home laboratory.

SCIENCECRAFT JUNIOR MICROSCOPE OUTFIT

Price $6.95 **Western Price $7.50**

The all metal microscope in this set is an unusually fine instrument to be sold at this price. The vials of specimen chemicals, glass slides, lens papers, and dissecting instruments are adequate equipment with which to perform all fundamental experiments and observations. The manual is complete and interesting. Packed in a heavy cardboard display box, size 15 x 10 inches.



THE MICROSCOPE

Most Important Tool in Science

Have you ever seen the barbs on a bee's stinger, the thousands of lenses in a fly's eye, or the hairs on an insect's leg? Did you ever watch crystals form from a drop of solution, or see wriggling vinegar eels that resemble snakes? With a ScienceCraft Microscope you can be an eyewitness to countless microscopic marvels of science. You will have experiences you will never forget, and you will discover secrets of nature you can learn in no other way.

All of the ScienceCraft Microscope outfits not only offer a splendid instrument to aid you in your study of the wonders of the world of microscopy, but all necessary accessories are also included. In the Senior, Master, and Master Deluxe you will find an exciting feature in that these sets contain three unique specimens—Uranium Ore, Gold Ore, and Silver Ore! This is the first time that these three unusual ores have ever been included in a Microscope outfit.

Science and the microscope go hand in hand. The microscope which Pasteur used is still the most important tool of science. Choose yours from the splendid assortment of ScienceCraft Microscope outfits. Velvet-action focusing adjustment, tilting frames, fully adjustable mirrors, and ground and polished lenses of high resolving power; make ScienceCraft Microscopes the leaders in the field.

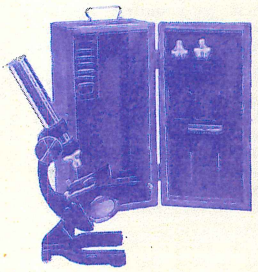
AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

ScienceCraft Microscope Accessories

By use of these accessories you will greatly increase the value and performance of your microscope outfit, as well as greatly increasing the number of observations and experiments that can be performed.

The Master Deluxe ScienceCraft Microscope Outfit already contains this special equipment. However, this list can of course be used in ordering replacements for this particular outfit. All other ScienceCraft Microscope Outfits can be built up through the separate purchase of any or all of these accessories.

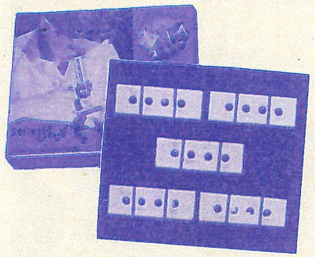
Low Power Objective (approximately 50 X)	$ .75
Medium Power Objective (approximately 100 X)	1.00
High Power Objective (approximately 200 X)	1.50
Better Vision Sub-Stage Illuminator with 2.2 Volt Magnifying Bulb	.50
Electronic Battery Holder (Batteries Not Included)	.35
Clear View Micro-Projector Attachment	.40
Micro-Illuminator for Opaque Subjects with two Magnifying Bulbs	.75
Electronic Stage	.35
Uranium Ore	.45
Silver Ore	.20
Gold Ore	.30
2.2 Volt Magnifying Bulb (For Sub-Stage Illuminator or Micro-Illuminator)	.15
Booklet, "How To Use Your ScienceCraft Microscope Attachments" Tells about many interesting pieces of equipment which can be added to your ScienceCraft Microscope.	.25
6—Sodium Ferrocyanide	.15
9—Ammonium Chloride	.10
10—Strontium Chloride	.20
14—Tartaric Acid	.15
26—Sodium Salicylate	.15
48—Mangnesium Sulfate	.10
57—Calcium Chloride	.10
61—Stearic Acid	.15
85—Urea	.15
Thinner	.15
Mounting Cement	.15
Glass Slides (Blank) per dozen	.20
Brush	.15
Tweezers	.15
Dissecting Needles per pair	.10
Rubber Bulb Pipette	.10
Glass Rod (5 mm. x 4")	.05
Lens Paper	.10
Blank Labels for Slides per 24	.10
Specimen Covers per dozen	.20
Diatomaceous Earth	.10
Sponge Skeleton	.10
ScienceCraft Senior Microscope	5.00
ScienceCraft Senior Microscopy Manual	.40
ScienceCraft Junior Microscope	2.50
ScienceCraft Junior Microscope Manual	.15



This microscope answers the demand for a separate instrument suitable for laboratory use by adults or advanced youths.

The microscope is all metal, with three interchangeable objective lens mountings of 50, 100, and over 150 powers. In small laboratory carrying case approximately 10½" inches high x 4¾" x 4". Complete with 6 glass slides and two dissecting needles. You will find this an exceptional value in an instrument of this type and quality.

STUDENT MICROSCOPE
No. 380 Price $12.95



Four-in-one set of mounted specimens on glass slides. There have been many requests for a set of mounted specimens for use with microscopes. This set is unusual in that each of the 5 slides contain four subjects, a total of 20 different specimens in all.

SLIDE SET
No. 351 Price $1.50

NOTE: Prices Shown Are 5% Higher In The 11 Western States

CHEMCRAFT REG US PAT OFF

ATOMIC ENERGY ACCESSORIES

The extremely important Atomic Energy feature is included in the No. 10, 15, and 25 Chemcraft Outfits, and the list of materials shown below will aid you in replacing these parts if you have used them up and require more.

If you are ordering this equipment for the first time, and not as a replacement we sincerely advise you to buy the COMPLETE ATOMIC ENERGY KIT—not only will you save money, but also, the various apparatus and equipment are of more value and interest to you when used together and as a series of experiments. For example, "The Story of Atomic Energy" manual is almost necessary to aid you in an understanding of how to use the Spinthariscope, Uranium Ore, Radio-active Screen, and Uranium Chemical.

Our supply of these kits is at present time limited because of the Government control on Uranium Ore, and other parts, but we will do our best to make immediate shipment to you.

PRICE LIST FOR ATOMIC ENERGY UNITS

"The Story of Atomic Energy" Manual	$ .75 each
Chemcraft Spinthariscope	.75 each
Uranium Ore	.45 each vial
Uranium Chemical	.45 each vial
Radioactive Screen	.45 each
	(Total $2.85)

Complete Atomic Energy Kit packed in Box $2.50

PLANTCRAFT SETS AND NUTRIENT CHEMICALS

FOR SOILLESS GARDENING

Plantcraft Sets are ideal for indoor gardening, as no soil is required, and the nutrient chemicals which are so clean and easy to work with contain all of the elements essential to the healthy growth of flowers, plants, and vegetables.

If you would like to start a new indoor garden you may order the complete Plantcraft Sets. If you already have an indoor garden you will find the separate nutrient chemicals of great value in encouraging growth.

These unique chemicals feature a perfectly balanced plant diet of nutritive elements, based on exhaustive tests and a series of experiments extending over many years. The synthetic plant hormone which supplements the natural hormones in all plants, produces healthier and more vigorous flower, vegetable, and plant development.

From the standpoint of economy Plantcraft Nutrient Chemicals give a great quantity of nutrient solution for little money, the cost being as low as two-thirds of a cent per gallon.

No. 603 Plantcraft Set, Complete	$3.50
No. 605 Plantcraft Set, Complete	5.95
Plantcraft Nutrient Chemicals, 2 oz. (1 oz. each A and B)	.40
1 lb. (8 oz. each A and B)	1.50
Plantcraft Instruction Folder (Included in Sets)	.20
BOOK NO. 1—Plant Propagation—Seeds, Cuttings, Hormones (including Propagation Chart and Cutting Propagation Chart)	1.50
BOOK NO. 2—Proper Conditions for Plant Growth, Soils, Fertilizers, Nutrients	1.00
BOOK NO. 3—Plant Diseases and Insects—Insecticides, Fungicides, Traps, Baits, Repellants	1.95
BOOK NO. 4—Plant Genetics—Artificial Pollenation and the Production of Hybrids	1.25

BRYAN VALENCE BLOCKS

PRESENT VISUALLY THE FUNDAMENTAL PRINCIPLES OF CHEMISTRY

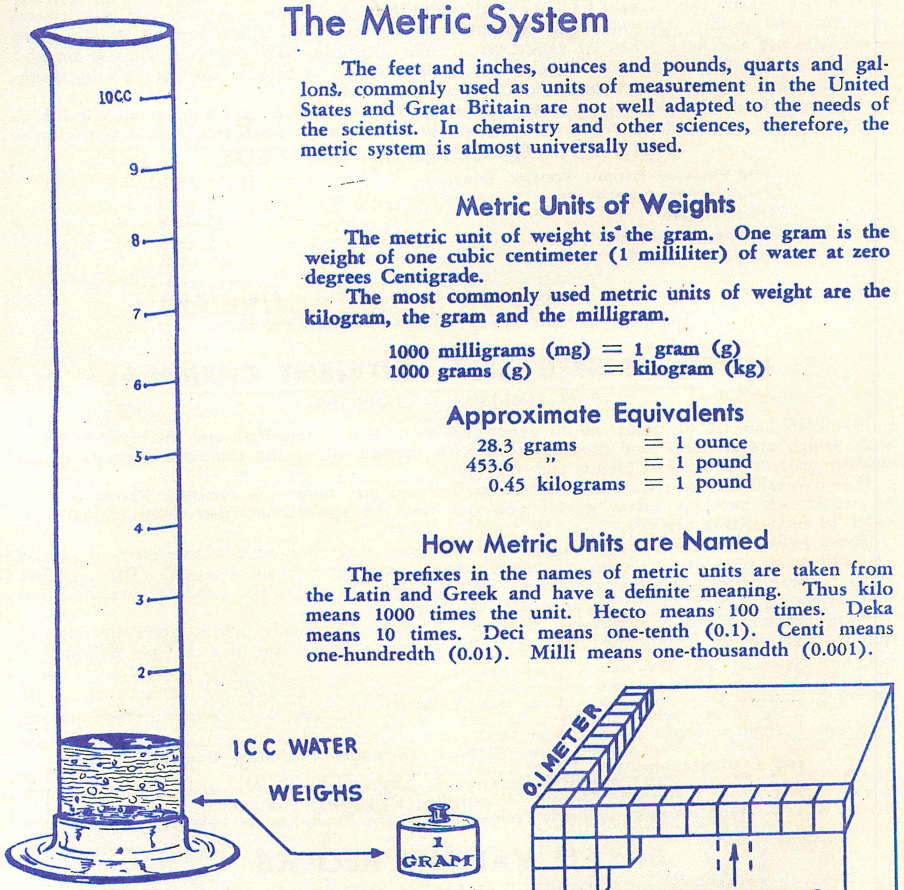
In order to fill the demand for a larger and more complete set of Chemical Illustrators, we offer the Bryan Valence Blocks which are specially designed for use in connection with first year chemistry classes.

The Bryan Valence Blocks consist of over 100 differently shaped pieces representing the more common chemical elements and molecules. These pieces are designed to fit together so as to illustrate correct chemical combinations. Each piece is lettered with the name, symbol and atomic weight of the element or molecule, which it represents. The pieces are distinctively colored, blue indicating metals and red indicating non-metals. In addition, the shade of the color indicates the relative chemical activity of the elements or molecules. Valence is visually represented by the number of arms extending from each piece or block. Full instructions for the use of the blocks accompanying each set.

The Bryan Valence Blocks are highly recommended by leading educators in the field of chemistry as a visual aid in helping pupils to a quicker and better understanding of the principles of theoretical chemistry. The blocks are packed in a convenient box with compartments for keeping the various sizes and colors separate. ----------PRICE $1.50

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

Useful Scientific Information



The Metric System

The feet and inches, ounces and pounds, quarts and gallons, commonly used as units of measurement in the United States and Great Britain are not well adapted to the needs of the scientist. In chemistry and other sciences, therefore, the metric system is almost universally used.

Metric Units of Weights

The metric unit of weight is the gram. One gram is the weight of one cubic centimeter (1 milliliter) of water at zero degrees Centigrade.

The most commonly used metric units of weight are the kilogram, the gram and the milligram.

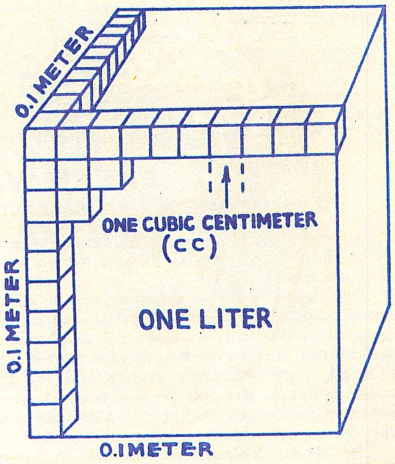
1000 milligrams (mg) = 1 gram (g)
1000 grams (g) = kilogram (kg)

Approximate Equivalents

28.3 grams = 1 ounce
453.6 " = 1 pound
0.45 kilograms = 1 pound

How Metric Units are Named

The prefixes in the names of metric units are taken from the Latin and Greek and have a definite meaning. Thus kilo means 1000 times the unit. Hecto means 100 times. Deka means 10 times. Deci means one-tenth (0.1). Centi means one-hundredth (0.01). Milli means one-thousandth (0.001).



Metric Units of Volume

The basic metric unit of volume is the liter, which is the volume of 1 kilogram of water at 4 degrees Centigrade. A cube with a volume of one liter measures almost exactly 0.1 meter (or 10 centimeters) on each edge. A liter is therefore equivalent to 10 x 10 x 10 (= 1000) cubic centimeters (cc).

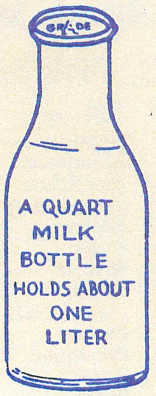
If a liter is divided into 1000 equal parts, each part is thus a cubic centimenter, or milliliter.

CHEMCRAFT REG U.S. PAT. OFF.

The commonly used metric units of volume are the liter and the milliliter. For all practical purposes the cubic centimeter and the milliliter are identical.

An ordinary teaspoon holds about 4 cc

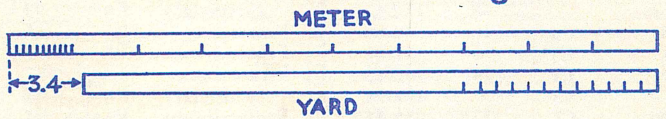
1000 milliliters (ml) = 1 liter (l)
1 milliliter = cubic centimeter (cc)



Approximate Equivalents

1 liter = 1.06 quarts
29½ milliliters = 1 fluid ounce
29½ cubic centimeters = 1 fluid ounce

Metric Units of Length

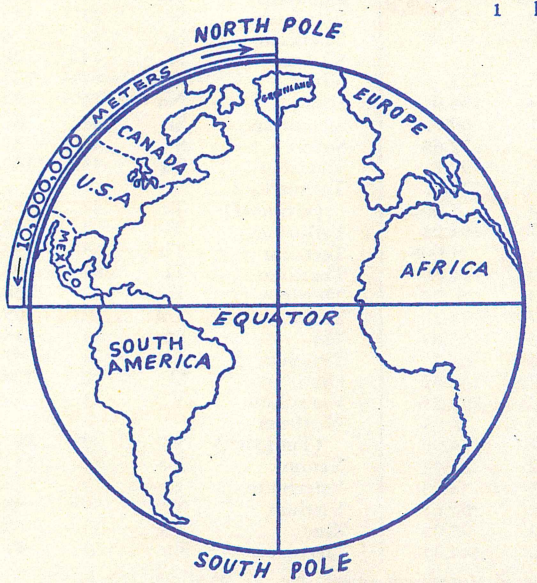


The most commonly used metric units of length are the kilometer, the meter, the centimeter and the millimeter.

10 millimeters (mm) = 1 centimeter (cm)
100 centimeters (cm) = 1 meter (m)
1000 meters (m) = 1 kilometer (km)

Approximate Equivalents

2½ centimeters = 1 inch
25 millimeters = 1 inch
1 meter = 39 inches
1 kilometer = 3/5 mile



The basic unit of length in the metric system is called the meter. A meter is approximately one-ten-millionth of the distance from the equator to the north pole. The meter is about 3 inches longer than the yard.

The meter is divided into 10 equal parts each of which is called a decimeter. The decimeter in turn is divided into 10 centimeters and the centimeter into 10 millimeters.

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

THE CHEMICAL ELEMENTS

Their Symbols, Atomic Numbers, and Atomic Weights

Element	Symbol	Atomic Number	Atomic Weight
Actinium	Ac	89	227.
Aluminum	Al	13	26.97
Americium (artificial)	Am	95	241.
Antimony	Sb	51	121.76
Argon	A	18	39.944
Arsenic	As	33	74.91
Astatine (artificial)	At	85	211.
Barium	Ba	56	137.36
Beryllium	Be	4	9.013
Bismuth	Bi	83	209.00
Boron	B	5	10.82
Bromine	Br	35	79.916
Cadmium	Cd	48	112.41
Calcium	Ca	20	40.08
Carbon	C	6	12.01
Cerium	Ce	58	140.13
Cesium	Cs	55	132.91
Chlorine	Cl	17	35.457
Chromium	Cr	24	52.01
Cobalt	Co	27	58.94
Copper	Cu	29	63.57
Curium (artificial)	Cm	96	242.
Dysprosium	Dy	66	162.46
Erbium	Er	68	167.2
Europium	Eu	63	152.0
Fluorine	F	9	19.00
Francium (artificial)	Fa	87	223.
Gadolinium	Gd	64	156.9
Gallium	Ga	31	69.72
Germanium	Ge	32	72.60
Gold	Au	79	197.2
Hafnium	Hf	72	178.6
Helium	He	2	4.003
Holmium	Ho	67	164.94
Hydrogen	H	1	1.008
Indium	In	49	114.76
Iodine	I	53	126.92
Iridium	Ir	77	193.1
Iron	Fe	26	55.85
Krypton	Kr	36	83.7
Lanthanum	La	57	138.92
Lead	Pb	82	207.21
Lithium	Li	3	6.94
Lutetium	Lu	71	174.99
Magnesium	Mg	12	24.32
Manganese	Mn	25	54.93
Mercury	Hg	80	200.61
Molybdenum	Mo	42	95.95
Neodymium	Nd	60	144.27
Neon	Ne	10	20.183
Nickel	Ni	28	58.69
Niobium (Columbium)	Cb	41	92.91
Neptunium (artificial)	Np	93	237.
Nitrogen	N	7	14.008
Osmium	Os	76	190.2
Oxygen	O	8	16.000
Palladium	Pd	46	106.7
Phosphorus	P	15	30.98
Platinum	Pt	78	195.23
Polonium	Po	84	210.
Plutonium (artificial)	Pu	94	239.
Potassium	K	19	39.096
Praseodymium	Pr	59	140.92
Promethium (artificial)	Pm	61	147.
Protactinium	Pa	91	231.
Radium	Ra	88	226.05
Radon	Rn	86	222.
Rhenium	Re	75	186.31
Rhodium	Rh	45	102.91
Rubidium	Rb	37	85.48
Ruthenium	Ru	44	101.7
Samarium	Sm	62	150.43
Scandium	Sc	21	45.10
Selenium	Se	34	78.96
Silicon	Si	14	28.06
Silver	Ag	47	107.880
Sodium	Na	11	22.997
Strontium	Sr	38	87.63
Sulfur	S	16	32.06
Tantalum	Ta	73	180.88
Technetium (artificial)	Tc	43	99.
Tellurium	Te	52	127.61
Terbium	Tb	65	159.2
Thallium	Tl	81	204.39
Thorium	Th	90	232.12
Thulium	Tm	69	169.4
Tin	Sn	50	118.70
Titanium	Ti	22	47.90
Uranium	U	92	238.07
Vanadium	V	23	50.95
Wolfram (Tungsten)	W	74	183.92
Xenon	Xe	54	131.3
Ytterbium	Yb	70	173.04
Yttrium	Y	39	88.92
Zinc	Zn	30	65.38
Zirconium	Zr	40	91.22

Note: The artificial production of elements of Atomic No. 97 Berkelium and Atomic No. 98 Californium was recently announced.

CHEMCRAFT REG U.S. PAT. OFF

Catalog of Refills, Apparatus, and Books

Refills may be purchased in 1 oz. Bottles, or in 4 oz. sturdy bulk containers. Filter papers, Test papers, etc., are of course furnished in Chemcraft Envelopes or Parcels.

	Refill	4oz.
1—Sulfur	$ .15	.35
2—Sodium Borate	.15	.35
3—Trisodium Phosphate	.15	.35
4—Sodium Carbonate	.15	.35
5—Sodium Bisulfite	.15	.35
6—Sodium Ferrocyanide	.20	.45
7—Sodium Bisulfate	.15	.35
8—Sodium Tungstate	.15	.55
9—Ammonium Chloride	.15	.35
10—Strontium Chloride	.20	.45
11—Zinc Metal	.10	.30
12—Aluminum Sulfate	.15	.35
13—Nickel Ammonium Sulfate	.15	.40
14—Tartaric Acid	.20	.45
15—Tannic Acid	.20	.45
16—Sodium Thiosulfate	.15	.35
17—Ferrous Ammonium Sulfate	.15	.40
18—Litmus Paper (State Color) 6 pc.	.15	
19—Sodium Silicate Solution (Water Glass)	.15	.35
20—Calcium Oxide	.15	.35
21—Ferric Ammonium Sulfate	.15	.40
22—Phenolphthalein Solution	.15	.35
23—Sodium Iodide Solution	.15	.35
24—Manganese Sulfate	.15	.35
25—Sodium Bicarbonate	.15	.35
26—Sodium Salicylate	.20	.45
27—Cobalt Chloride	.15	.35
28—Chameleon Paper	.10	
29—Logwood	.15	.35
30—Mixed Dyes	.15	.40
31—Sulfide Test Paper	.10	
32—Congo Red Paper	.10	
33—Glycerine	.15	.35
34—Blue Print Paper	.10	
35—Tumeric Paper	.10	
36—Powdered Iron Metal	.15	.35
37—Tin Metal	.20	.50
38—Calcium Hypochlorite	.15	.35
39—Azurite	.15	.35
40—Ammonium Molybdate	.35	.90
41—Starch	.15	.35
42—Phenolphthalein Paper	.10	
43—Iron Sulfide	.15	.35
44—Iron Pyrites	.15	.35
45—Chrome Alum	.15	.35
46—Ammonium Hydroxide	.15	.35
47—Conchineal	.15	.35
48—Magnesium Sulfate	.15	.35
49—Calcium Sulfate (Gypsum)	.15	.35
50—Calcium Carbonate (Precipitated Chalk)	.15	.35
52—Gum Arabic	.15	.35
54—Boric Acid	.15	.35
55—Powdered Charcoal	.15	.35
56—Charcoal Lump	.15	
57—Calcium Chloride	.15	.35
58—Flame Test Wire	.10	
59—Blank Test Paper	.10	
60—Filter Paper	.10	
61—Stearic Acid	.15	.35
62—Benzoic Acid	.15	.35
63—Graphite	.15	.35
64—Sodium Ammonium Hydrogen Phosphate	.15	.35
66—Bismuth Subnitrate Solution	.15	.35
67—Potassium Carbonate	.15	.35
68—Dilute Hydrochloric Acid 5%	.15	.35
70—Dilute Acetic Acid	.15	.35
75—Dimenthylglyoxime Solution	.15	.35

APPARATUS AND BOOKS

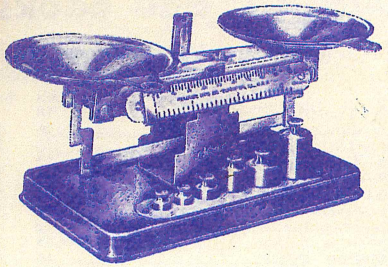
Streak Plate	.15
Glass Tubing, 11 x $\frac{4}{10}$ in. size—6 pcs.	.25
Hydroquinone	.15
Potassium and Sodium Tartrate	.15
Saccharine Solution	.15
Muscovite	.10
Selenite	.10
Zinnwaldite	.10
Rose Quartz	.10
Resin	.10
Willemite	.15
Wernerite	.15
Graphite	.10
Franklinite	.15
ScienceCraft Micro-Projector Bulb	.50
Hematoxylin Stain	.15
Eosine Stain	.15
Methyl Green Stain	.15
Methyl Violet Stain	.15
Glass Rod (5 mm. x 4")	.05
Lens Paper	.10
Eight-hole Test Tube Rack (without drying pegs)	.35
Watch Glass	.20
Mortar and Pestle (1 ounce)	.50
Flask (Florence) 60 ml	.40
Flask (Florence) 250 ml.	.60
Flask (Erlenmeyer) 125 ml.	.50
Beaker (Resistant Glass) 50 ml.	.40
Beaker (Resistant Glass) 100 ml.	.50
Funnel	.35
Byran Chemical Illustrators	.75
Gas Generating Apparatus Complete	.50
Gas Generating Apparatus without Funnel	.25
Byran Valence Blocks	1.50
Gas Delivery Tube and Stopper	.10
Wash Bottle Assembly Complete (250 ml. Flask, Stopper, 3 pieces of bent Glass Tubing)	.90
Glass Tube	.05
Assorted Glass Tubing (7 pieces)	.25
Test Tubes (4 x ½ inch size) 5c ea. 40c per dozen	
Test Tubes (5 x ⅝ inch size) 10c ea., 3 for 25c	
Porcelain Crucible, 15 ml. No. 0	.85
Chemcraft Alcohol Lamp Blow Torch	.50
Alcohol Lamp (1 oz.)	.25
Alcohol Lamp Wick	.05
Hard Wax Candle	.05
Stirring Rod, Spoon, and Measure	.15
Metal Chemcraft Balance	1.25

Books and Instruction Manuals

The Chemcraft Beginners Manual (1951 Edition)	.50
The Chemcraft Junior Manual (1951 Edition)	.75
The Chemcraft Senior Manual (1951 Edition)	1.25
The Chemcraft Master Manual	1.50
Chemcraft Chemical Magic Manual	.35
Chemcraft Glass Blowing and Construction of Laboratory Apparatus Manual	.35
Chemcraft Qualitative Analysis Manual	.35
ScienceCraft Micro-Projector Manual	.25
Chemcraft Rubber Chemistry Manual	.50
ScienceCraft Senior Microscope Manual	.40
ScienceCraft Mineralogy Manual	1.00
Microscope Attachment Manual	.25

See page 5 on "How to Order"

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS



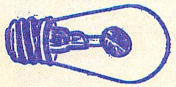
UNUSUAL NEW OFFERINGS

No. 200 NEW SPECIAL PELOUZE SCALE

This is the finest small professional laboratory Scale available. It is the latest model, improved over previous Pelouze Scales. Nine new features: all metal, accurate to less than one grain, six brass metric weights, knife-edge pivot, self-locating sliding poise, dual readings in metric and avoirdupois systems. Often compared with the $20.00 Scale. Price each $8.95

No. 201 ASSORTMENT OF RANDOM LENGTH ALNICO V MAGNETS

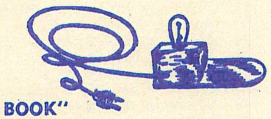
Available only through this catalog. Small sections (about ¼ to 1½ inch). Strongest Alnico permanent magnets. About 20 to an assortment. Regular value if not broken lots would be at least $7.00. Price each Assortment only $2.00.



No. 202 "BLACK LIGHT" ARGON LAMP

Glow lamp for generation of ultra-violet and black light. Fits standard socket for 110 volts AC or DC. Price each 75c

No. 203 MICROSCOPE AC SUB-STAGE ILLUMINATOR



A highly important accessory for improved illumination — fits all Sciencecraft Microscopes, and most others. Price each, complete with cord $1.45

No. 204 "THE COLLECTOR'S HAND BOOK"

An interesting and stimulating new booklet covering a wide variety of the most popular collector's subjects. Of equal value to the hobbyist and home scientist. Please do not order until January 1, 1951. Price each ---------- $1.00.

ELEMENTS OF CHEMISTRY by Brownlee and four other authors.

This is one of the best elementary text books. It covers the field of inorganic chemistry thoroughly and touches here and there on the simpler forms of organic chemistry. Many schools use this book. Its 684 pages are profusely illustrated. Cloth bound. Price ---------- $3.00

LABORATORY EXERCISES IN CHEMISTRY by Brownlee and other authors.

This book is written to accompany the above text book. It gives nearly a hundred different experiments. Contains much information about the care of the apparatus and chemical manipulation. Over 300 pages helpfully illustrated. Price ------ $1.50

HOW TO MAKE AND USE A SMALL CHEMICAL LABORATORY
By Raymond Francis Yates
Just the Book You Have Been Looking For.
Tells what you need to equip your laboratory.
How to use and take care of your equipment.
How to make some of your apparatus yourself.
How to perform many excellent experiments.
Filled from Cover to Cover With Chemical Facts You Want to Know. Price ---------- $1.00

Henley's Twentieth Century Book of FORMULAS, RECIPES AND PROCESSES
Latest Edition
One Useful Recipe Will be Worth More Than Ten Times the Price of the Book
Here is a book of 800 pages that is the most complete Book of Formulas ever published, giving thousands of recipes for the manufacture of valuable articles for every-day use. Hints, Helps, Practical Ideas and Secret Processes are revealed within its pages. It covers every branch of the useful arts and tells thousands of ways of making money and is just the book everyone should have at his command. A book to which you may turn with confidence. A mine of information, up-to-date in every respect. Price ---------- $4.00

CATALOG OF SPECIAL CHEMICALS

The chemicals listed below are packed in attractive containers neatly labeled with name, quantity and formula. We have listed each chemical under the name most frequently used, and have also given the formula and other common names.

NO.	NAME, FORMULA AND PHYSICAL DESCRIPTION	1 oz.	4 oz.
305	ACID ACETIC 36% (Solution) $CH_3.COOH$ (Vinegar Acid)	$0.15	.35
315	ACID HYDROCHLORIC dilute 5% (Solution) HCl (Hydrogen Chloride) (Muriatic Acid) POISON	.15	.35
318	ACID SALICYLIC $OH.C_6H_4CO_2H$ Needlelike Crystals	.15	.35
322	ACID SULFURIC Dilute 5% (Solution) H_2SO_4 (Hydrogen Sulfate) (Oil of Vitriol) POISON	.20	.50
335	ALLOY FUSIBLE (Wood's Metal)	.35	1.00
345	ALUMINUM POTASSIUM SULFATE $Al_2K_2(SO_4)_4.24H_2O$ White powder—clear irregular Crystals (Potash Alum) (Ordinary Alum)		.20
401	BARIUM CARBONATE $BaCO_3$ White Powder POISON	.15	.35
402	BARIUM CHLORIDE $BaCl_2.2H_2O$ Flat Colorless Crystals POISON	.15	.35
406	BARIUM SULFATE $BaSO_4$ Heavy White Powder (Barytes) (Heavyspar) POISON	.15	.35

See page 5 on "How to Order"

CHEMCRAFT REG U.S. PAT. OFF

512	CALCIUM CARBONATE $CaCO_3$ Grayish White Lumps (Marble Chips)	.20
514	CALCIUM FLUORIDE CaF_2 Grayish White Powder (Fluorspar)	.20
515	CALCIUM HYDROXIDE Ca $(OH)_2$ Fine White Powder (Calcium Hydrate) (Slaked Lime)	.20
519	CALCIUM PHOSPHATE—Monobasic $CaH_4(PO_4)_2.H_2O$ White Powder (Monocalcium Phosphate)	.15 .35
522	CALCIUM SULFIDE CaS Grayish or Yellowish White Powder (Luminous Calcium Sulfide)	.25 .65
533	CHROMIUM POTASSIUM SULFATE $Cr_2K_2(SO_4)_4 24H_2O$ Violet Crystals (Chrome Alum) POISON	.15 .35
567	COPPER CARBONATE (Basic) $Cu_2(OH)_2CO_3$ Green Powder (Cupric Carbonate) (Artificial Malachite) POISON	.15 .35
568	COPPER CHLORIDE $CuCl_2.2H_2O$ Green Needle-like Crystals (Cupric Chloride) POISON	.15 .35
571	COPPER OXIDE CuO Black Powder (Cupric Oxide) (Black Copper Oxide)	.15 .35
573	COPPER SULFATE $CuSO_4.5H_2O$ Blue Crystals (Blue Stone) (Blue Vitriol) POISON	.15 .35
591	DEXTRINE Yellow or White Powder (Starch Gum) (British Gum)	.10 .25
801	FEHLING'S SOLUTION Alkaline Solution or Copper Solution	.20 .50 .20 .50
922	GUM CAMPHOR $C_9H_{14}CO$ White Translucent mass (Camphor)	.25 .65
1080	HYDROGEN PEROXIDE H_2O_2 3% Solution	.30
1161	IRON (ic) AMMONIUM SULFATE $Fe_2(NH_4)_2(SO_4)_4.24H_2O$ Light Orange Crystals	.15 .35
1163	IRON (ic) CHLORIDE $FeCl_3.6H_2O$ Yellowish Crystaline Mass (Ferric Chloride)	.15 .35
1165	IRON (ic) FERROCYANIDE $Fe_4(Fe(CN)_6)_3$ Dark Blue Powder (Ferriferrocyanide) (soluble Prussian Blue) POISON	.25 .65
1166	IRON (ic) OXIDE Fe_2O_3 Reddish Brown Powder (Ferric Oxide) (Brown Oxide of Iron	.15 .35
1168	IRON (ous) SULFATE $FeSO_4.7H_2O$ Bluish green crystals (Copperas) (Ferrous Sulfate) (Green Vitriol)	.15 .35
1420	LEAD METAL—(Granular) Pb Bluish Gray coarse powder	.25
1426	LEAD OXIDE Pb_3O_4 Bright Red Heavy Powder (Red Lead) (Minium) (Red Oxide of Lead) POISON	.15 .35
1512	MAGNESIUM CHLORIDE $MgCl_2.6H_2O$ White Crystals	.15 .35
1515	MAGNESIUM OXIDE MgO Fine White Powder (Magnesia)	.20 .50
1531	MANGANESE DIOXIDE MnO2 Black POWDER (Manganese Peroxide) (Pyrolusite)	.15 .35
1555	METHYL ORANGE INDICATOR $(Ch_3)_2N.C_6H_4.N_2.C_6H_4.SO_2Na$ Orange Colored powder	.60 1.45
1625	NICKEL NITRATE $Ni(NO_3)_2.6H_2O$ Emerald Green Crystals	.25 .65
1627	NICKEL SULFATE $NiSO_4 7H_2O$ Green Crystals (Nickel Vitriol)	.15 .35
1810	PHENOLPHTHALEIN $(OHC_6H_4)_2CO.C_6H_4CO$ White Powder	.25 .65
1811	PHENOLPHTHALEIN SOLUTION IN ALCOHOL—INFLAMMABLE	.15 .35
1821	POTASSIUM BITARTRATE $KHC_4H_4O_6$ White Powder (Cream of Tartar)	.15 .35
1822	POTASSIUM BROMIDE KBr White Crystalline granules	.15 .35
1825	POTASSIUM CHLORIDE KCl White Crystals (Muriate of Potash)	.15 .35
1829	POTASSIUM FERROCYANIDE $K_4Fe(CN)_6.3H_2O$ Yellow Crystals (Yellow prussiate of Potash)	.20 .50
1838	POTASSIUM SULFATE K_2SO_4 Small Colorless Crystals or powder	.15 .35
1850	PARAFFIN Colorless wax-like substance	.20
2129	SODIUM ACETATE $NaC_2 H_3O_2.3H_2O$ Colorless Crystals	.15 .35
2132	SODIUM BISULFATE $NaHSO_4$ White Crystals (Sodium Acid Sulfate) (Nitre Cake) POISON	.10 .25
2137	SODIUM CARBONATE $Na_2CO_3.10H_2O$ Colorless Crystals (Soda) (Sal Soda) (Washing Soda)	.35
2143	SODIUM HYDROXIDE NaOH Hard White Pellets or Sticks (Soda Lye) (Sodium Hydrate) (Caustic Soda) POISON	.15 .35
2144	SODIUM IODIDE NaI.$2H_2O$ White Granular Powder	.25 .65
2147	SODIUM NITRITE $NaNO_2$ Colorless Crystals	.15 .35
2149	SODIUM PHOSPHATE, Tribasic $Na_3Po_4.12H_2O$ Crystals or White powder (Trisodium Phosphate (t. s. p.)	.10 .25
2150	SODIUM PHOSPHATE, DIBASIC $Na_2HPO_4.H_2O$ White Powder	.15 .35
2151	SODIUM PHOSPHATE, MONOBASIC $NaH_2PO_4.H_2O$ Transparent Crystals (Sodium Acid Phosphate) (Monosodium Phosphate)	.20 .50
2154	SODIUM SULFATE Na_2SO_4 White Crystals (Salt Cake)	.10 .25
2155	SODIUM SULFITE Na_2SO_3 White Powder	.10 .25
2230	TIN (ous) CHLORIDE $SnCl_2.2H_2O$ White Crystalline Mass (Stannous Chloride) (Tin Salt)	.25 .65
2280	TURMERIC Yellowish Powder (Curcumin)	.15 .35
2821	ZINC METAL, SHEETS Zn Grayish metal	.10 .25
2823	ZINC CHLORIDE $ZnCl_2$ Grayish White Powder POISON	.15 .35

See page 5 on "How to Order"

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

Catalog of Minerals and Supplies

Collecting, Classifying and Studying Minerals is a profitable and enjoyable hobby for everyone

ScienceCraft Mineralogy Manual $1.00

This 150 page manual has been approved and endorsed by many authorities as the most practical book on Mineralogy for amateurs. There are chapters on geology, chemistry, crystallography, physical properties of minerals, and blow pipe and chemical tests. Under the heading "Descriptive Mineralogy" the properties, descriptions, locations, uses and complete physical and chemical tests for more than 125 native American minerals are given. Following this are chapters on Gems, How to Collect Minerals, and finally tables for quick determination of various minerals.

Mineralogy Testing Equipment and Reagents

NO.					
2	Sodium Borate	$ .15	Alcohol Lamp Complete (1 oz.)	.25
4	Sodium Carbonate	.15	Alcohol Lamp Complete (2 oz.)	.50
11	Zinc Metal (granular)	.15	Wick for Alcohol Lamp	.05
27	Cobalt Chloride Solution	.15	Streak Plate	.15
31	Sulfide Test Paper	.15	Test tubes (4 x ½" size) 5c each, 40c per doz.	
55	Powdered Charcoal	.15	Glass Tubing (11 x 7/16" dia.) 10 pcs.	.40
56	Charcoal Block	.15	Measure	.05
58	Flame Test Wire	.10	Magnet	.10
68	5% Hydrochloric Acid (1 oz. bot.)	.15	Steel Mortar	.50
69	Manganese Dioxide	.15	Watch Glass (3")	.20
	Hard Wax Candle	.05		

LIST OF MINERAL SPECIMENS

The Mineral Specimens in the following list have been collected from famous localities in the United States and many other countries.

These minerals are ¾" size or larger. Each specimen is numbered in accordance with the Dana System.

Price each $0.15

No.	Mineral	Location
2	Graphite	Ceylon
2A	Graphite in Rock	New York
3	Native Sulfur	Utah
13	Gold Bearing Quartz	Virginia
13A	Gold Ore	Maryland
14	Silver Ore	Arizona
15	Copper (Native)	Pennsylvania
25A	Iridescent Iron Ore	South Dakota
45	Galena	Oklahoma
45A	Galena on Fluorite	Nova Scotia
54	Chalcocite	Montana
58	Sphalerite	Oklahoma
66	Cinnabar	Arkansas
74	Pyrrhotite	Canada
78	Bornite	Montana
83	Massive Chalopyrite	Colorado
83A	Chalcopyrite and Bornite	Colorado
83B	Massive Chalcopyrite & Magnetite	Pennsylvania
83C	Chalcopyrite on Dolomite-Xls.	Oklahoma
85	Massive Pyrite	Missouri
85A	Pyrite Cubes	Virginia
85B	Pyrite in Quartz	Virginia
85C	Pyrite Cubes in Chlorite Schist	Vermont
85D	Nodular Pyrite	Indiana
96	Marcasite	Oklahoma
98	Arsenopyrite	New York
166	Halite Xls. in Clusters	California
167	Sylvite and Halite—Xline.	New Mexico
175	Fluorite	Nova Scotia
*175B	Fluorescent Brown Fluorite on Celestite	Oklahoma
183	Cryolite	Greenland
210	Rock Crystals	Arkansas
210A	Glassy Quartz	Maryland
210D	Rose Quartz	South Dakota
210E	Amathestine Quartz	Virginia
210F	Smoky Quartz	Virginia
210G	Milky Quartz	Pennsylvania
210N	Gray Quarzite	South Dakota
210O	Red Quartzite	South Dakota
210P	Brown Quartzite	South Dakota
210R	Agatized Wood	South Dakota
210V	Petrified Wood	Oklahoma
210W	Flint	Denmark
210X	Gray Jasper	South Dakota
210Z	Red Jasper	South Dakota
212D	Common Opal	Nevada
212E	Milk Opal	Nebraska
*212G	Fluorescent Opal	Nebraska
212O	Diatomaceous Earth	California
212R	Tripolite	Oklahoma
224	Cuprite	Pennsylvania
228	Zincite	New Jersey
231	Corundum Crystals	South Africa
231B	Corundum—massive, pure	South Africa
231C	Massive Emery	New York
231D	Sapphire Rock	Utah
232	Hematite—massive	Maryland
232A	Specular Hematite	Pennsylvania
232C	Red Paint Ore	South Dakota
234	Spinel	New Jersey
236	Gahnite	New Jesrey
237A	Magnetite—pure variety	New York
239	Franklinite	New Jersey
*239A	Franklinite Xls. on Fluorescent Calcite	New Jersey
241	Chromite	Maryland
243	Hausmannite	Washington
248	Cassiterite	South Dakota
250	Rutile Ore	Maryland
254	Pyrolusite	Africa
257	Goethite	Pennsylvania
258A	Dendrites	Pennsylvania
259	Ocherous Limonite (Ocher)	Maryland
259A	Stalactitic Limonite	Pennsylvania
259B	Limonite Pseudomorph after Pyrite (cubes)	Maryland
259D	Limonite	Maryland
259E	Velvety Limonite	North Carolina
261	Bauxite	Arkansas
269	Psilomelane	Virginia
270	Dog Tooth Spar Crystals	Oklahoma
270A	Clevable Calcite	Virginia

See page 5 on "How to Order"

CHEMCRAFT REG U.S. PAT. OFF.

No.	Mineral	Location
270B	Orange Calcite with Pyroxene	New York
270C	Orange Calcite	Maryland
270F	Fossil Limestone	Pennsylvania
270G	Black Marble	Virginia
270H	Chalk	England
*270L	Granular Fluorescent Calcite	New Jersey
270M	Pink Granular Calcite	Maryland
270N	Mottled Limestone	South Dakota
270P	Calcite with Pholgopite	Maryland
271	Pink Dolomite	Oklahoma
271A	White Dolomite	New York
271B	Cleavable Dolomite	Maryland
271C	Red Dolomite on Serpentine	Pennsylvania
272	Magnesite	Maryland
273	Siderite	Connecticut
274	Rhodochrosite	Montana
277	Aragonite—Flos Ferri Variety	South Dakota
280A	Massive Strontianite in Calcite	Maryland
289	Azurite	South Dakota
313	Orthoclase Feldspar Crystals	Nevada
313D	Pink Feldspar	South Dakota
313E	White Feldspar	South Dakota
315	Amazonstone	Virginia
315B	Yellow Feldspar	Virginia
315C	Ordinary Feldspar	Virginia
315E	Crystallized Feldspar	New Jersey
315F	Orange Feldspar and Hematite Xls. in Quartz	Pennsylvania
316	Albite	Virginia
316A	Clevelandite	Virginia
317	Oligoclase	North Carolina
323	Enstatite	North Carolina
323A	Bronzite	North Carolina
325A	Diopside	Maine
325B	Coccolite in Calcite	New York
325C	Hedenbergite	Montanna
327	Spodumene	South Dakota
330	Pectolite	New Jersey
*335	Rhodonite	New Jersey
335A	Fowlerite	New Jersey
338F	Green Actinolite	Pennsylvania
338G	Actinolite on Hornblende	North Carolina
338J	Amphibole Asbestos	North Carolina
338K	Mountain Leather	Maryland
338P	Hornblende	North Carolina
338X	Tremolite in Dolomite	Connecticut
344B	Beryl	South Dakota
344C	Brown Beryl	New York
357	Nephelite	Canada
360	Cancrinite	Maine
362	Sodalite	Canada
370A	Grossularite Garnet	Connecticut
370E	Iron Garnet	South Dakota
370G	Garnet Xls. in Schist	South Dakota
376	Olivine	North Carolina
*381	Fluorescent Willemite	New Jersey
*387	Fluorescent Wernerite (Scapolite)	Canada
393	Vesuvianite	Maine
397	Massive Topaz	Virginia
397A	Topaz Crystal	Utah
398	Andalusite	South Dakota
399	Sillimanite (with Quartz)	South Dakota
400	Kyanite—Crystalline	Virginia
406	Zoisite	Massachusetts
407A	Amygdaloid Epidote in Altered Rhyolite	Pennsylvania
408	Piedmontite	Pennsylvania
409	Allanite—radio-active	North Carolina
411	Prehnite	New Jersey
415	Chondrodite in Rock	New York
423	Calamine	Oklahoma
426	Massive Tourmaline	Virginia
426A	Tourmaline Xls. in Clusters	South Dakota
426B	Tourmaline Xls. on Mica in Gneiss	Connecticut
426C	Blue Tourmaline	South Dakota
426D	Pink Tourmaline	Maine
427	Dumortierite	California
428	Staurolites in Chloritoid	Maine
458	Muscovite Mica	Virginia
458B	Sericite	California
458C	Silver Mica Schist	South Dakota
458D	Yellow Mica Schist	South Dakota
460	Lepidolite	South Dakota
460A	Amber Lepidolite	South Dakota
460B	Cookeite	Connecticut
461	Zinnwaldite	Virginia
462	Biotite	North Carolina
462A	Phlogopite	New Jersey
462C	Biotite with Quartz and Feldspar	South Dakota
462D	Eastonite	Pennsylvania
462E	Caswellite	New Jersey
462F	Biotite Schist	Maryland
466A	Chlorite in Matrix	North Carolina
466B	Chlorite in Quartz	Pennsylvania
468	Clinochlore Schist	New York
468A	Penninite on Basalt	New Jersey
469	Crystallized Prochlorite	Maryland
480	Zonolite	Montana
480A	Vermiculite (scales)	North Carolina
481	Serpentine (semi-precious)	Pennsylvania
481E	Chrysotile Asbestos	Arizona
481 F	Serpentine—1 polished face	Maryland
481G	Asbestos on Tremolite and Serpentine	Pennsylvania
481J	Picrolite	Maryland
482	Deweylite on Magnesite	Maryland
484B	Talc—Soapstone Blocks	Maryland
484D	Massive Talc	Maryland
484F	Foliated Talc	Maryland
490A	Glauconitic Sand (Greensand)	Virginia
492	Kaolin	Virginia
496	Montmorillonite with Manganapatite	Maine
497	Pyrophyllite (white)	North Carolina
497A	Red Pyrophyllite	North Carolina
502A	Thaumasite with Strontianite	New Jersey
522	Microlite Xls. on Matrix	Virginia
544	Lithiophyllite	Connecticut
549A	Apatite	Canada
549B	Green Granular Apatite	New Jersey
549C	White Apatite with Smoky Quartz	New Jersey
549E	White Apatite on Matrix	Maryland
552	Vanadinite	Arizona
555A	Griphite	South Dakota
559	Amblygonite	South Dakota
574	Lazulite in Pyrophyllite	North Carolina
639	Wavellite	Arkansas
642	Turquoise Matrix	California
664	Carnotite—radio-active	Utah
701	Howlite	California
704	Colemanite	California
711A	Xls. of Thorianite on Matrix	Pennsylvania
719	Barite	West Virginia
719A	"Desert Rose" Barite	Oklahoma
719B	Nodular Barite	Oklahoma
720A	Celestite and Fluorite	Ohio
746	Rock Gypsum	Virginia
746A	Selenite Gypsum—Xline	Oklahoma
746B	Satin Spar Gypsum	South Dakota
825	Anthracite	Pennsylvania
826	Oil Shale	Tennessee
828	Bituminous Coal showing Color Interference	Pennsylvania
828D	Lignite	Montana
830	Pumice	Italy
831	Obsidian	California
832	Bituminous Sandstone	California
833	Umber	Italy

FLUORESCENT MINERALS

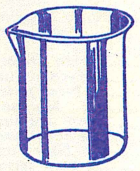
Minerals which show the beautiful phenomenon of fluorescence under the argon bulb, iron arc light, or mercury vapor lamp, are marked with an asterisk (*).

See page 5 on "How to Order"

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

Chemical Laboratory Equipment

No. 1—ASBESTOS BOARD, 4" square, for use as a protection when heating Beakers, Evaporating dishes flasks, etc.
Price Each $0.15



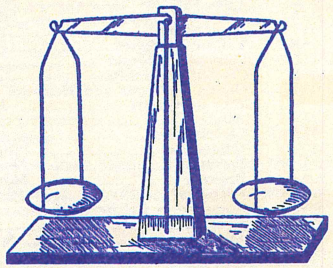
No. 34—BEAKERS

A few beakers in your laboratory will be a great convenience. They can be used for heating liquids, whereas a tumbler would break as soon as heat was applied. Our beakers are made of best quality resistant glass and will be found highly satisfactory.

Capacity	50ML	100ML	150ML	250ML
Price each	$0.40	$0.50	$0.50	$0.60

No. 16—CHEMICAL BALANCE

This balance is suitable for weighing small quantities of chemicals up to 10 grams. It is sensitive to 0.1 grams (one three-hundreth of an ounce). Height: 5½ inches. Two 1 gram weights are included with each balance.
Price each $1.25



No. 17—Additional 1 gram weights $0.15 each.

No. 47—HOMEOPATHIC VIALS

Homeopathic vials with lip finish.

Size	2 dram
Cork Size	No. 1
Price each	$0.05
Price per dozen	0.40

REAGENT BOTTLES FOR EVERY PURPOSE

The best containers for laboratory chemicals and reagents are standard bottles. The bottles listed below are the best obtainable for this purpose and the different styles should be chosen according to the kind of reagent you wish to keep in them.

For storing acids and other liquids which would be liable to attack corks, and for voliatile liquids, use narrow mouth glass stoppered bottles.

For solids which are affected by air or which give off corrosive fumes, use wide mouth glass stoppered bottles.

For ordinary liquids, use narrow mouth cork stoppered bottles.

For ordinary solids, use wide mouth cork stoppered bottles.

The stoppers in our glass stoppered bottles are ground to fit the bottles perfectly and the proper size corks and rubber stoppers for cork stoppered bottles are noted, but they are not included with the bottle.

No. 41—BOTTLES—NARROW MOUTH WITH GLASS STOPPERS
(Tinctures)

Capacity	1 oz.	2 oz.	4 oz.	8 oz.
Height	3¾ in.	4¾ in.	5½ in.	7 in.
Diameter	1¼ in.	1⅜ in.	1⅞ in.	2⅜ in.
Price each	$0.35	$0.40	$0.45	$0.55

No. 42—BOTTLES—NARROW MOUTHS FOR CORKS

Capacity	1 oz.	2 oz.	4 oz.	8 oz.
Cork Size	No. 2	No. 3	No. 4	No. 5
Rubber Stopper Size	No. 00	No. 0	No. 0	No. 1
Price each	$0.05	$0.08	$0.12	$0.18
Price per doz.	$0.50	$0.80	$1.20	$1.80

See page 5 on "How to Order"

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No. 43—BOTTLES—WIDE MOUTH WITH GLASS STOPPER
(Salt Mouths)

Capacity	1 oz.	2 oz.	4 oz.	8 oz
Price each	$0.35	$0.40	$0.45	$0.55

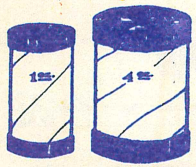
No. 44—BOTTLES—WIDE MOUTH FOR CORKS

Capacity	1 oz.	2 oz.	4 oz.	8 oz.
Cork Size	No. 8	No. 10	1¼ in.	1⅜ in.
Rubber Stopper Size	No. 3	No. 4	No. 6	No. 7
Price each	$0.05	$0.08	$0.12	$0.18
Price per dozen	.50	$0.80	$1.20	$1.80

No. 46—SHELL VIALS

The shell vials are the same as those used in CHEMCRAFT and make convenient containers for many purposes. Takes No. 4 corks.

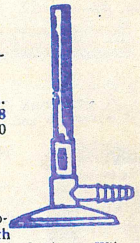
Size	4 x ½ in. flat bottom
Price each	$0.05
Price per dozen	.40



No. 52—BOXES (Pasteboard)

Pasteboard Boxes make excellent containers for solid chemicals which are not affected by the air.

Capacity	1 oz.	2 oz.	4 oz.	8 oz.
Price each	$0.05	$0.06	$0.07	$0.08
Price per dozen	0.35	0.40	0.50	0.60



No. 55—BUNSEN BURNER

In this burner, devised by Robert Bunsen a jet of ordinary illuminating gas is projected from a narrow opening into a wider tube. In this tube the gas becomes mixed with air, drawn in through openings whose dimensions can be altered by means of a perforated ring. When the supply of air is sufficient, the flame becomes non-illuminuous and intensely hot.
Price each $0.80

No. 56—WING TOP FOR BUNSEN BURNER

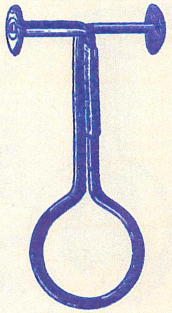
The wing top may be put on or taken off a Bunsen Burner easily and quickly. It makes a fan-shaped flame, which enables you to bend glass tubing more satisfactorily than with the regular flame as it heats the glass for several inches instead of in only one spot.
Price each $0.20

No. 60—SPECIAL RUBBER TUBING FOR BUNSEN BURNER

Per foot $0.10

No. 66—CANDLES

Hard wax patent candles for laboratory use.
Price each $0.05



No. 69—CLAMP (Burette)

The most satisfactory and generally used clamp. Ideal for attaching and holding a burette, condenser, flask or test tube to an upright such as a ring stand. Adjustable to any angle.
Price $0.75

No. 71—CLAMP (Pinchcock)

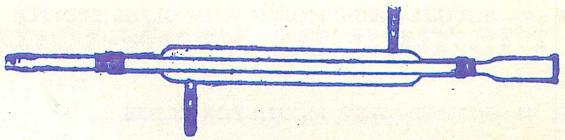
This clamp is for use with rubber tubing and is very easy to use because it stops the flow as soon as released by the fingers. With these clamps you can place tight valves throughout your apparatus without the expense of purchasing glass stop cocks.
Price each $0.25

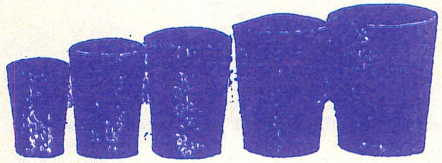
See page 5 on "How to Order"

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

No. 75—CONDENSER (Liebig's)

The Liebig condenser is designed expressly for the purpose of converting hot vapors to liquids. The hot vapors pass through the central condenser tube, which is covered with a jacket in which circulating water keeps the central tube cool, causing the vapors to condense. The liquids can be collected in a beaker or flask. Liebig's condenser, length of jacket 10 inches. Price each ---------- $1.75





No. 78—CORKS (Regular Length—Taper)

We supply only the best grade of corks. Sizes to fit any bottle.

Cork Size No.	0	1	2	3	4	5	6	7	8	9	10	11	12
Diameter in in.	¼	⅜	½	9/16	⅝	11/16	¾	13/16	⅞	15/16	1	1 1/16	1⅛
Price each	1c	1c	2c	2c	2c	3c	3c	3c	3c	4c	4c	4c	4c
Price per 10	10c	10c	15c	15c	15c	25c	25c	25c	25c	30c	30c	30c	30c

No. 80—CORK BORER

A good cork borer is useful in any laboratory. This one bores six sizes of holes, making a clean cut and is easy to use. Cork borer set of 1-6. Price each ---------- $2.80

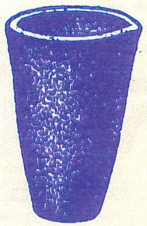
No. 83—CRUCIBLES (Porcelain) High Form

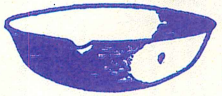
For experiments which require the heating of solids, melting metals, etc., crucibles are a necessity. Porcelain crucibles are best suited for fusing small quantities of materials over a bunsen burner or alcohol lamp, especially when the materials are of such a nature as to attack iron. Porcelain crucibles with cover.

	No. 0	No. 1	No. 2
Capacity	15ML	30ML	50ML
Price each	$0.85	$1.20	$1.45

No. 84—CRUCIBLES (Clay) Size D

Clay crucibles round form with lip pouring. Can be used for melting metals in your furnace or stove. Capacity about 6 oz. 4" high, 2⅜" top diameter.
Price each ---------- $0.35





No. 90—EVAPORATING DISHES

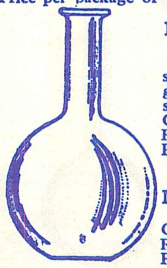
A good evaporating dish is essential for evaporating solutions and for carry out reactions that necessitate heating liquids or pastes. Best quality porcelain evaporating dish.

Capacity	70ML	90ML
Price	$0.70	$1.10

No. 94—FILTER PAPER

Best quality laboratory filter paper, gives clear filtrate. 2½ inch size fits 1½ inch funnel, 5 inch size fits 3 inch funnel.

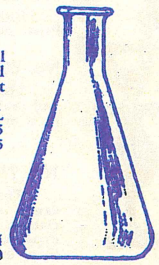
Diameter	2½ in	5 in
Price per dozen sheets	$0.05	$0.08
Price per package of 100 sheets	0.15	0.25



No. 101—FLASKS (Florence) (Round-Flat Bottom)

There are many uses for flasks in the laboratory and several sizes should be kept on hand. The durable flasks we list will give the best service. The rubber stoppers which fit the different sizes are noted below:

Capacity	60ML	125ML	250ML	500ML
Rubber Stopper Size	NO. 2	No. 3	No. 4	No. 5
Price each	$0.40	$0.50	$0.60	$0.75



No. 102—FLASKS (Erlenmeyer)

Capacity 250ML
Rubber stopper size ---------- No. 4
Price each ---------- $0.60

See page 5 on "How to Order"

CHEMCRAFT REG U.S. PAT. OFF

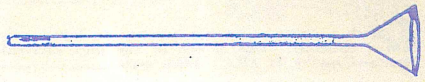
No. 112—GAUZE

In heating flasks, beakers, evaporating dishes, etc., it is very important to protect the bottom of the vessel by using wire gauze. Our wire gauze is fine mesh and has an extra protection of asbestos woven in the center.
Wire gauze with asbestos center 4 inches square ---------- $0.20

No. 106—FUNNELS

Good funnels are necessary in every laboratory not only for use in filtering, but also for filling narrow mouth bottles without spilling the liquids.
Best glass funnels, diameter 1¼ inch. Price each ---------- $0.35
Best glass funnels, diameter 3 inch. Price each ---------- $0.70

No. 108—FUNNEL TUBE



A funnel tube is very necessary for introducing liquids into a gas generating flask, as it makes it unnecessary to remove the cork from the flask and prevents air from becoming mixed with the gas. A funnel tube also acts as a safety valve and removes all danger of bursting apparatus.
Best quality glass funnel tube, length about 10 inches. Price each ---------- $0.45

No. 115—GLASS ROD—3/16

Our glass rod is the best obtainable. You will find a few pieces very handy to have on hand for making stirring rods of various length, etc. Furnished in eleven inch length.
Diameter 3/16 inch. Price per 10 pieces ---------- $0.35

No. 116—GLASS TUBING

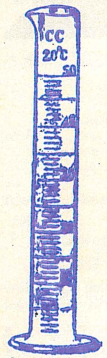
Every laboratory should have an assortment of various sizes of glass tubing. There is hardly a piece of apparatus that you may want to make that will not require some glass tubing. With a little practice it is very easy to cut and bend glass tubing.

In breaking a glass tube first make a scratch with a three cornered file, extending about half way around the tube, then break the tube in your fingers or across a sharp edge as you would break a stick. Always put the strain on the tube at the point where the scratch is made.

To bend a glass tube place the part where the curve is required in the flame of an alcohol lamp or bunsen burner, using a wing top if possible. The whole of the surface must be heated and the tube should be revolved in the flame.

Larger tubing is much harder to break and bend than the smaller sizes. Best grade glass tubing in eleven inch length. No. 3 wall.

Outside diameter	1/8 in.	5/16 in.	¼ in.	½ in.
Price, per 10 pieces	$0.30	$0.35	$0.45	$0.65



No. 118—GRADUATE (Cylinder)

You will need a good graduate for measuring liquids. Our 50 ml graduate with pouring lip has wide base so that it stands firmly.
Graduated in milliliters from 1 to 50.
Capacity 50ML. Price each ---------- $1.50

Capacity 2 ounces.

No. 123—JAR (Hydrometer)

This type of jar is designed especially for use in taking the hydrometer readings of liquids. The jar is filled with the liquid and the hydrometer is then floated inside being plainly visible through the clear glass. This jar is also very useful for other purposes. Size 1½ inches in diameter by 10 inches high. Price each ---------- $0.80

No. 126—LABELS

All of your laboratory chemicals should be labeled so as to avoid mistakes. We furnish neat blank labels, well gummed which will add greatly to the appearance of your reagent shelves.
Gummed labels—per pack of 100 ---------- $0.35

No. 128—LAMP (Alcohol) 2 oz.

If it is impossible to have a bunsen burner in your laboratory, a good alcohol lamp will serve as a substitute. Our alcohol lamps are made of heavy glass with cap to prevent evaporation of alcohol when not in use. Complete with wick.
Price ---------- $0.50

See page 5 on "How to Order"

AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

No. 141—MORTAR AND PESTLE

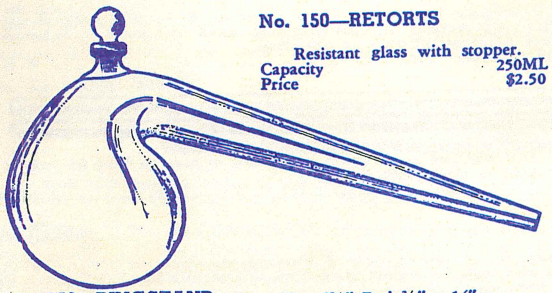
Many experiments require the pulverizing of chemicals in order to aid their reaction or mixture with other substances. For this purpose a glass mortar and pestle is used. Our mortars and pestles are made of heavy glass and the mortars have lips for pouring.

Mortar and Pestle, capacity 1 oz. Price each ---------- $0.50
Mortar and Pestle, capacity 4 oz. Price each ---------- $1.00

No. 145—PIPETTE

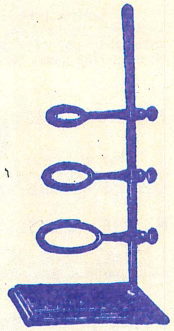
With a pipette a known quantity of liquid may be drawn from a bottle and safely transferred to any desired container. This is one of the most convenient and accurate methods of measuring liquids.
Pipette—Capacity 25 ML. Price each ---------- $1.10

No. 150—RETORTS



Resistant glass with stopper.

Capacity	250ML	500ML
Price	$2.50	$3.25



No. 152—RINGSTAND—Base 4" x 6¾" Rod ⅜" x 14".

A good ringstand when properly fitted up will be used more than almost any other piece of apparatus. Our ringstands are best grade, oak base with iron rod.
Ringstand without rings. Price each ---------- $0.75

No. 152A—RINGSTAND—Base 5¼ x 7¼ Rod ⅜ x 20 inches.

This larger ringstand is of heavier construction than No. 152, having a larger base of cast iron.
Ringstand without rings. ---------- $1.60

No. 153—RINGS FOR RINGSTANDS

These cast iron rings clamp securely onto our ringstands. Easily adjustable to any height and are held by a winged screw attachment.

Diameter	2 inch	3 inch	5 inch
Price each	$0.70	$0.80	$1.10

No. 156—RUBBER STOPPERS

Rubber stoppers seal tighter than corks and are less easily attacked by acids. We furnish the best grade soft rubber stoppers, either solid or with one or two holes. These rubber stoppers will be found very convenient when setting up apparatus. Specify whether solid, 1 hole or 2 hole.

Size No.	00	0	1	2	3	4	5	6	7	8
Diameter in inches	⅜	7/16	¾	13/16	15/16	1 1/16	1⅛	1¼	1 7/16	1⅝
Price each	.02	.03	.04	.05	.06	.07	.08	.12	.15	.20
Price per Doz.	.20	.25	.35	.45	.55	.65	.75	1.00	1.25	1.50

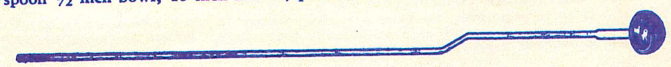
No. 157—SYNTHETIC TUBING

This tubing is especially made for chemical laboratory use. It is resistant to chemicals and will give excellent service.

Inside diameter	⅛ in	3/16 in.	¼ in.
Price, per foot	$0.05	$0.07	$0.09

No. 164—SPOON (Deflagrating)

Frequently it is desired to burn a solid or liquid in a bottle or vessel containing a certain gas. In this case a deflagrating spoon is necessary and will be found a very useful article.
Deflagrating spoon ½ inch bowl, 16 inch handle, price each ---------- $0.35



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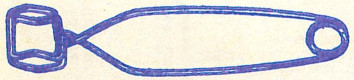
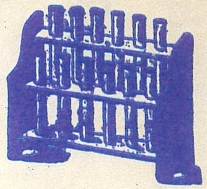


No. 173 TEST TUBE HOLDER

Do not burn your fingers. Use a test tube holder. Made of spring steel. Length over all 4 inches.
Price each ---------- $0.10

No. 174—TEST TUBE RACK



Wood test tube rack with holes for standing tubes up-right and pegs for draining washed tubes. You should have one in your laboratory.
Price each ---------- $0.70

No. 175—Small Test Tube Rack

Metal test tube rack, eight holes, without draining pegs.
Price each ---------- $0.35

No. 170—TEST TUBES

High quality test tubes made by automatic machines from flint glass tubing. Uniform walls and a well formed lip. Each tube annealed and free from stains.

Size inches	4x½	5x⅝	6x¾	7x⅞
Price each	$0.05	$0.10	$0.12	$0.15
Price per doz.	.40	.75	.90	1.00

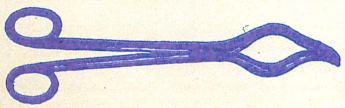
No. 172—TEST TUBE BRUSH

Always keep your test tube clean. Better results will be obtained. These brushes give a firm scrubbing surface from top to, and including the rounded bottom of test tubes.

Price each ---------- $0.15

No. 178—THERMOMETER

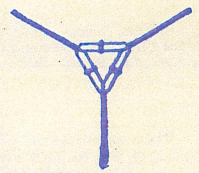
Chemical thermometer graduated with both Centigrade and Fahrenheit scales. An excellent thermometer for any laboratory.
Price each ---------- $2.25



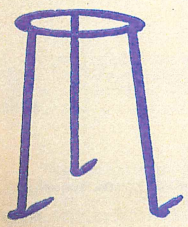
No. 180—TONGS (Crucible)

A good pair of crucible tongs come in very handy in case you have anything hot to pick up or hold.
Tong-steel. Price each $0.60

No. 183—TRIANGLES



Iron wire triangles covered with clay pipestems make good supports to use on ringstands or tripods to hold crucibles, etc., when heating. Length of sides 2 inches. Iron wire pipestem triangles.
Price each ---------- $0.25



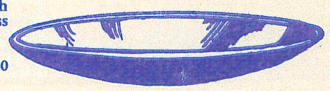
No. 185—TRIPOD

A good iron tripod is very convenient for heating flasks, beakers, evaporating dishes, etc.

Small Tripod	Height 6 in.	Price $0.50
Large Tripod	Height 9 in.	Price $0.75

No. 194—WATCH GLASS

For small experiments in crystallization, evaporating volatile liquids, etc., watch glasses are very convenient. Our watch glasses are made of well annealed glass with smooth ground edges.
Diameter about 3 inch.
Price each ---------- $0.20



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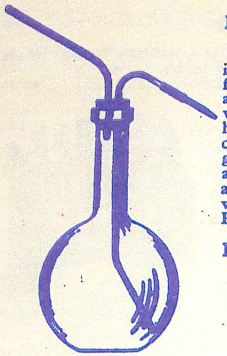
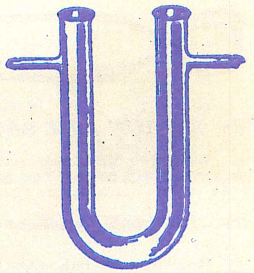


No. 191—U TUBES

These tubes are used for drying or washing gases. If you wish to remove moisture from the gas, put calcium chloride in the tube and as the gas passes through it the moisture will be absorbed. The gas may be washed by having water or other liquid in the tube and causing gas to bubble through. U tubes are generally connected to the gas generator and after passing through them the gas is collected and used as desired. Best grade glass U tube with side tubes, height 4 inches.
Price each ---------- $0.70



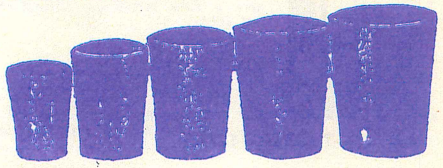
No. 197—WASHING BOTTLES

With these bottles you can direct a fine jet of water into a test tube, flask or anywhere you desire. In a laboratory where running water is not available, they are practically indispensable.
Capacity 250 Ml. Price each ---------- $0.90

Apparatus Assortment

Look over these assortments carefully, you will find many bargains among them. We cannot break an assortment. If you do not want the assortment just as we have listed it, order your apparatus from the individual list. Order these assortments by number and include the required postage.

Assortment No. 4500
CORKS



Includes each size from number 1 to number 6. The sizes you use most frequently. About 50 corks altogether. Price ---------- $0.40

Assortment No. 4505
RUBBER STOPPERS

Rubber stoppers are in constant use in any laboratory. Here is an assortment of popular sizes that will save your money.

	Solid	1 hole	2 hole
Size 0, diameter of large end 11/16 inch	1	1	1
Size 1, diameter of large end 3/4 inch	1	1	1
Size 2, diameter of large end 7/8 inch	1	1	1
Size 3, diameter of large end 15/16 inch	1	1	1
Size 4, diameter of large end 1 1/16 inch	1	1	1
Size 5, diameter of large end 1 1/8 inch	1	1	1

18 Stoppers altogether. Special Assortment Price ---------- $0.90

Assortment No. 4510
GLASS TUBING AND ROD

Fifteen pieces assorted glass tubing and rod. Tubing, 1/8, 11/16, 1/4, 1/2, inch diameter, Rod 1/16 inch diameter. This assortment will give you just the glass tubing and rod you need for your laboratory.
Special Assortment Price ---------- $0.45

Assortment No. 4515
SYNTHETIC TUBING

Every experimenter needs flexible tubing. This assortment is just what you want.
3 feet 1/8 inch inside diameter Synthetic Tubing
3 feet 1/16 inch inside diameter Synthetic Tubing
3 feet 1/4 inch inside diameter Synthetic Tubing
Special Assortment Price ---------- $0.50

Assortment No. 4520
FUNNEL AND FILTER PAPER

This assortment will take care of your filtering problems in fine shape. Includes a 3 inch funnel and a pack of filter paper to fit.
Price ---------- $0.80

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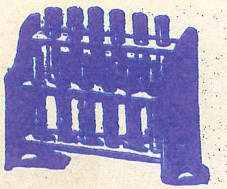


Assortment No. 4530
TEST TUBES

Four of each of the most popular sizes of test tubes 4x1/2, 5x3/8, 6x3/4 inches. One dozen altogether. This assortment will be welcomed by every experimenter.
Price ---------- $0.75

Assortment No. 4535
TEST TUBES AND TEST TUBE RACK

This assortment includes the same test tubes as the assortment 4530 and in addition, a test tube rack as shown in the illustration.
Special Assortment Price ---------- $1.25

Assortment No. 4536
TEST TUBES AND SMALL TEST TUBE RACK

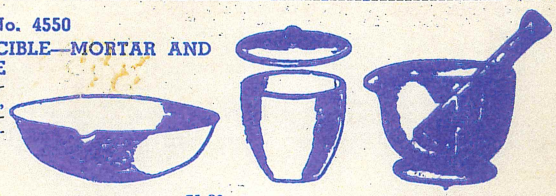
This assortment consists of four test tube 4x1/2, four test tubes 5x3/8 inches and a small wood test tube rack. Special Assortment Price ---------- $0.75

Assortment No. 4540
FLASK AND BEAKERS

One 50ML beaker, one 60ML flask, one 150ML beaker and one 125 ML flask. Four articles, the most popular sizes. Made of special resistant glass.
Price ---------- $1.50

Assortment No. 4550
EVAPORATING DISH—CRUCIBLE—MORTAR AND PESTLE

This assortment includes one porcelain evaporating dish, capacity 2 oz., one porcelain crucible with cover, capacity 30ML, one mortar and pestle, capacity 1 oz.



Price ---------- $1.90

Assortment No. 4560
BUNSEN BURNER COMPLETE

Here is an outfit every laboratory should have. A bunsen burner fitted with a removable wing top and 24 inches of best quality tubing all ready to connect to your gas supply.
Price ---------- $1.00

Assortment No. 4570
BUNSEN BURNER AND LARGE TRIPOD

Includes a bunsen burner-two feet tubing-iron tripod and wire gauze; a mighty handy assortment.
Price ---------- $1.55

Assortment No. 4575
LARGE TRIPOD AND WIRE GAUZE

An iron tripod is very convenient for heating flasks; beakers, and evaporating dishes, and the asbestos centered wire gauze protects them from breaking. Can be used with either an alcohol lamp or bunsen burner. Special Assortment Price ---------- $0.95

Assortment No. 4580
LARGE RINGSTAND COMPLETE WITH BUNSEN BURNER

Here is a dandy. One (No. 152A) adjustable ringstand. Three adjustable rings different sizes. One adjustable clamp for ringstand. One wire gauze. One bunsen burner with 24 inches best quality tubing.
Price ---------- $4.50

Assortment No. 4585
LARGE RINGSTAND WITH RINGS AND CLAMP

This includes an adjustable ringstand with three sizes of rings, and an adjustable clamp for holding flasks or burette. One of the most useful pieces of equipment you can add to your laboratory.
Special Assortment Price ---------- $3.25

Assortment No. 4600
DISTILLATION APPARATUS

This assortment will answer two purposes. First—for distillation the log delivery tube should be wrapped with a cloth and kept damp with cold water to hasten condensation. Second—it may be used as a gas generator when heat is required to give off a gas. Included in this assortment are: one 16-oz. flask, one two-hole rubber stopper to fit, one funnel tube, one long delivery tube with rubber connection.
Price ---------- $0.95

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AMERICA'S ORIGINAL AND LEADING CHEMISTRY OUTFITS

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HAND BOOK AND CATALOG

1951

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THE GREAT UNKNOWN

Fifty years ago it was evident that X-rays were useful for seeing through objects, such as the human body, which are opaque to ordinary light. It could not be predicted that X-rays would become a powerful weapon in the fight against cancer, or that the study of X-rays would lead to the discovery of radioactivity and stimulate the development of radio, radar, and a host of electronic devices. Such unforseen developments are the result of every great discovery. It would, therefore, be surprising if the most important consequences of the release of atomic energy are not in directions as yet unpredictable.

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TO THE MYSTERIES AND MIRACLES
OF AN ATOMIC WORLD

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FOREWORD

The authentic information on atomic energy presented in this book has been gathered from a large number of reliable sources. Included among these sources are the works of noted physicists and chemists and the latest reports released by the United States Atomic Energy Commission.

At the time of going to press there is not, to our knowledge, any other popular book covering the subject of atomic energy as complete as herein set forth.

THE STORY OF ATOMIC ENERGY WITH SAFE EXPERIMENTS

by
Harold M. Porter
and
Jermain D. Porter, PhD.

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The Porter Chemical Company
Hagerstown, Maryland

INTRODUCTION

Every new discovery in science adds new revelations and new problems to our already complicated civilization. People of more advanced years often look back to the days of their youth when life was much more simple and responsibilities fewer. We can not ignore scientific progress, however, and it is the duty of all of us to welcome new discoveries in a spirit of progressiveness and to equip ourselves mentally to do our part in guiding these new discoveries into constructive instead of destructive applications.

Thus it is with atomic energy. Guided into channels of usefulness, atomic energy will become as great a servant to mankind as fire. Allowed to follow destructive channels, atomic energy could conceivably destroy our world. Controlled it can become the most beneficial force ever discovered by man.

The basis of civilization is knowledge. Therefore, let us learn about atomic energy and each be prepared to do his part in making this great discovery benefit all mankind. It is with this thought that we present this elementary booklet describing atomic energy to you — the youth of America.

We have endeavored to make the explanation of atomic energy relatively simple and easily understandable, and for this reason, we have not gone into the many complicated theories and conplex mathematical equations which are involved in a more thorough presentation of the subject. We have presented experiments that you can perform yourself, and which will illustrate many of the interesting theories described in the booklet. With this elementary knowledge of atomic energy, you will be in position to understand future developments and to do your part as time goes on to help atomic energy make our world a better place in which to live.

H. M. P.

CHAPTER I

HISTORY OF ATOMIC ENERGY

For more than a century, since John Dalton proposed his atomic theory, scientists have known that all matter is composed of atoms, but it is only much more recently that we have even suspected what a tremendous store of energy lies locked within each atom. And only within the past few years have we learned how a portion of this energy can be released from certain atoms.

John Dalton

Let us go back half a century, to 1896. X-rays had just been discovered, and scientists were busily looking for other examples of invisible rays. Henri Becquerel in Paris had been experimenting with a compound containing uranium and one day happened to place some of it in a drawer with some photographic plates. The plates, to his amazement, became fogged, even though they had been wrapped in black paper. Further experiments showed that the uranium compound was giving out invisible but penetrating rays, rays which passed through black paper and even through thin sheets of metal.

Pierre and Marie Curie, also working in Paris, at once began to investigate this remarkable phenomenon. They soon discovered that pitchblend, an ore of uranium emitted stronger radiation than uranium itself. After six years of the most difficult and discouraging labor, they obtained, starting from tons of pitchblende, a tiny quantity of the first pure salt of radium, a hitherto unknown metal, which sends out rays a million times as intense as those from uranium.

Henri Becquerel

Within a few years is had been established that radioactivity, this peculiar property possessed by uranium, radium, and a few other substances is due to an actual breaking down, or disintregration, of their atoms.

The study of radioactivity was one of the chief sources of our present view of the atom as being composed of a nucleus, containing positively charged protons, surrounded by negative electrons.

In 1919, Lord Rutherford, using rays from a radioactive substance, changed artificially for the first time one kind of atom into another, on, however, an extremely small scale.

Another important step based upon the use of radioactivity as a research tool was taken in 1932, when Chadwick in England discovered the neutron, a particle without electrical charge. Neutrons are now known to be present along with protons in the nuclei of atoms. They play a very important part in the release of atomic energy.

Pierre and Marie Curie

With the development of the cyclotron and similar "atom-smashing" devices around 1932, it became possible to bombard atoms with a greater variety of particles having very high energies. The research worker was no longer dependent upon expensive radioactive materials for his high-speed particles, but could produce them at will in great abundance.

With these new resources it was very quickly proved that the famous Einstein mass-energy equation was correct. This equation, based upon the Einstein theory of relativity, tells us how much energy is obtained when a given weight of matter is converted into energy.

At about the same time, Fermi in Rome was systematically studying the effect of neutrons upon each of the chemical elements. When he reached the heaviest kind of atom then known, that of uranium, he obtained some curious results. These were not correctly explained until 1939, when Hahn and Strassmann proved that bombardment of uranium by neutrons causes the uranium atoms to split apart into two more or less equal fragments. This was an entirely new finding, totally different from ordinary radioactivity. As we shall see later, this splitting, or atomic fission, is accompanied by the release of a tremendous amount of energy. It is the key process in the production of atomic energy.

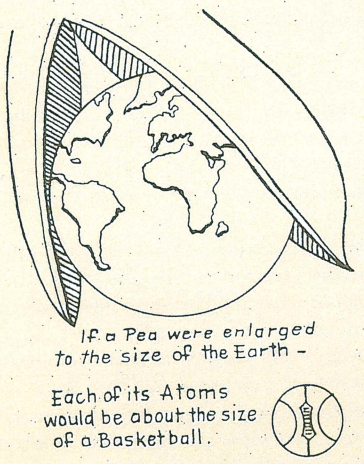
Between 1939 and the first explosion of an atomic bomb in 1945, there was feverish activity among the atomic scientists, most of it under the cloud of war censorship. It was found that only one rather rare kind of uranium atom (U-235) is subject to fission and that when it splits, several neutrons are released, which may then cause fission of another atom, and so on. Thus the possibility of a self-sustaining chain reaction was recognized. Further, the entirely new element, plutonium, was made by the action of neutrons upon ordinary uranium atoms (U-238) and was shown to undergo fission itself.

Finally, a large number of very formidable scientific and engineering problems were overcome during the war years to bring into practical application the laboratory experiments and the theoretical results we have just sketched. The climax was the atomic bomb, but that was not the end, for with the coming of peace, the atomic scientists have applied their talents to the problem of constructive uses for atomic energy, and the history of a new era is just beginning to be written in their laboratories.

CHAPTER II
STRUCTURE OF THE ATOM

All substances — elements, compounds, and mixtures — are composed of atoms. Thus atoms are the building blocks of all matter, whether in the solid, liquid, or gaseous state, including air, water, rocks, metals, everything we eat, drink, and wear, and even our own bodies.

It used to be thought that atoms were hard little balls quite impossible to split up into smaller pieces. We know now, however, that atoms, tiny though they are, are themselves built up of still smaller units. In this chapter we are going to learn something about these units and how they are combined into atoms. It is only by understanding the structure of atoms that we can tap the tremendous supplies of energy within them.

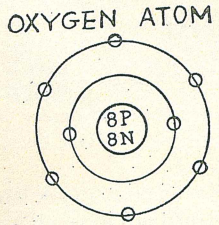


Even the largest of the atoms is so inconceivably small that it would mean little to say that it measures only a few billionths of an inch across. Perhaps it may help to put it another way and state that there are 30 billion times more atoms in the head of a pin than there are people on the earth.

Or, from still another point of view, suppose we imagine a pea enlarged until it is the size of the earth. Then each of the atoms which make it up, enlarged in the same proportion, would still be no larger than a basketball.

From the point of view of chemical behavior nearly a hundred different kinds of atoms are now known, corresponding to the ninety some known elements. If we consider, however, atoms which differ merely in weight but not in chemical nature, the number of different kinds amounts to several hundred. Two atoms of the same chemical element which differ from each other only in weight are said to be *isotopes* of that element. For example, tin has ten stable isotopes, for there are ten different kinds of tin atoms all mixed together in a piece of ordinary tin. They all behave chemically like tin atoms, but differ among themselves in weight.

Because of their extremely small size, atoms cannot be seen even with the most powerful microscope. Nevertheless, from experiments which tell us about their behavior, chemists and physicists have been able to deduce many things about their structure. We know with certainty that nearly all of the mass, or weight, of an atom, more than 99.9% of it at least, is concentrated at its very center, in a very small fraction of the atom's volume. We call this heavy little core the *nucleus*. Locked within it is the greater part of the atom's store of energy.



Moving about the nucleus at fairly definite distances are tiny particles bearing charges of negative electricity. These are called *electrons*. Compared with the size of the nucleus and of the electrons, the distances between them are very large, and we must imagine the atom as being composed mostly of absolutely empty space.

If we could magnify a tiny atom to the size of a great football stadium, the nucleus might be represented by a golf-ball lying on the 50-yard line, while the electrons would be like a swarm of mosquitos buzzing around the outer row of seats. Or the atom may be compared with our solar system, in which the planets, among them our earth, revolve at enormous distances about the sun.

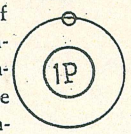
Since the electrons all have charges of negative electricity, they repel each other. It is the repulsion between the electrons of neighboring atoms which makes it impossible for two objects to occupy the same space at the same time.

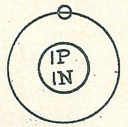
For an understanding of atomic energy, the nucleus is more important than the electrons. The nucleus of atoms is composed of two kinds of particles, *protons* and *neutrons*. Protons are positively charged particles, about 1840 times as heavy as electrons. Neutrons are particles of about the same weight as protons, but they have no electrical charge.

The number of protons in the nucleus of an atom is the *atomic number* of that atom. Since protons and electrons have equal, though opposite, electrical charges, there must be as many electrons around the nucleus of a normal neutral atom as there are protons in the nucleus. Then the negative electrons will exactly counterbalance the positive charge of the protons, and the atom as a whole will be neutral.

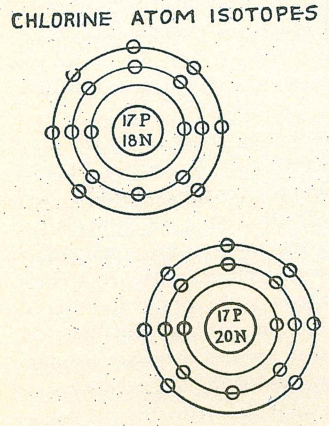
The total number of particles in the nucleus — protons and neutrons together — is called the *mass number* of the atom. It corresponds closely to the atomic weight used by the chemist. Since the protons and neutrons are equal in weight, while the electrons are negligible in comparison, the mass numbers of the atoms tell us their relative weights, just as the atomic numbers tell us the number of units of positive charge on their nuclei.

All the isotopes of the same element have the same atomic number, but differ in mass number, that is, in atomic weight. This is the same as saying that they have the same number of protons in their nuclei, but different numbers of neutrons. If the number of protons is the same for two or more isotopes, then the number of electrons will also be the same. Now in chemical reactions, only the outer electrons are concerned, so that two atoms with the same number of electrons will show identical chemical behavior, and the chemist will say they belong to the same element. This agrees with what we said above about isotopes being identical chemically.





Perhaps some of these things will be made clearer by a few examples. Consider first the simplest of all atoms, that of ordinary hydrogen. Its atomic number and its mass number are both one (1), so its nucleus consists of a single proton, and of course, there is a single electron revolving around the nucleus. Hydrogen has an isotope, the so-called heavy hydrogen, or deuterium. Its atomic number is naturally, also one, but its mass number is two, showing its nucleus to contain besides a proton, also one neutron. Thus the only difference between these two isotopes of hydrogen is that the heavier one has a neutron in its nucleus.



Ordinary chlorine gas is a mixture of two isotopes, both having the atomic number 17, meaning 17 protons in the nucleus. One has a mass number of 35, showing that its nucleus has 18 neutrons in

addition to the 17 protons (to make a total of 35 particles). The mass number of the other is 37, and its nucleus therefore is made up of 20 neutrons plus 17 protons.

The "average mass number" of the isotopes in ordinary chlorine is 35.46, and is what the chemist called the atomic weight of chlorine. It is closer to 35 than to 37 because the isotope of mass number 35 is present in greater amount.

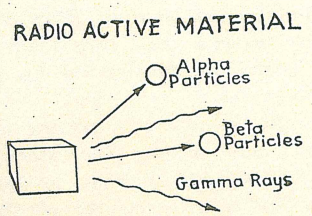
In the next chapter we are going to see that an understanding of isotopes is essential in obtaining a clear conception of radioactivity.

CHAPTER III
RADIOACTIVE ATOMS

In the nucleus of an ordinary atom the protons and neutrons appear to be cemented together into a balanced and stable structure. Atoms of a few of the less common elements, however, have nuclei which are so unstable that they break down spontaneously with a sort of nuclear explosion. Such atoms are said to be radioactive, and as pointed out in Chapter I, the study of radioactivity was one of the important steps toward the mastery of atomic energy.

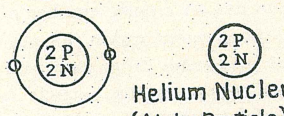
A sample of a radioactive element steadily gives off energy in the form of radiation and in doing so changes into other elements. Nothing we can do has the slightest effect upon the rate at which this change proceeds. We can chill it below the temperature of liquid air or bake it at white heat of the electric furnace, but the disintegration of the atoms and the emission of energy go right on at exactly the same speed. We can join the atoms into chemical compounds or put them under tremendous pressure, still without the slightest effect upon their decomposition.

Evidently we are dealing here with something much different from the usual chemical reactions which atoms undergo. The carbon of coal, for example, combines rapidly with oxygen (burns) at a red heat, but at room temperature the combination does not take place at all. In a chemical reaction such as the burning of carbon, atoms of two different elements (carbon and oxygen) combine to form molecules of a compound (carbon dioxide) with the release of



energy in the form of heat. Radioactivity, on the other hand, is an actual decomposition of single atoms. A radioactive atom shoots out a ray of some sort and thereafter is an atom of an entirely different element. The total energy released is thousands of times greater than that from the combustion of an equal weight of coal.

Three types of rays are emitted by radioactive substances. They are called alpha, beta, and gamma rays.



Alpha rays are streams of rapidly moving particles which are the same in mass and charge as the nuclei of helium atoms. That is, they have a mass number of 4 and bear two units of positive charge. The two positive charges show that the alpha particle contains two protons, and mass number indicates the presence in them of two neutrons in addition. Alpha particles, although they are shot out with speeds of around 10,000 miles per second, travel only a short distance before being stopped by collisions with atoms in their path. They are stopped by even a thin sheet of paper or metal foil, but in air they may travel as far as an inch or two.

Beta rays are streams of electrons, moving with tremendous speeds — up to 184,000 miles per second. They are much more penetrating than the slower-moving alpha particles and will pass through thin sheets of metal. Heavier metals like lead are more effective in stopping them than light metals such as aluminum.

Gamma rays are not particles at all, but a sort of wave motion like X-rays, though for the most part of shorter wave length and greater penetrating power. They are not completely stopped even by several inches of lead. Since they are not particles, they have no mass and carry no electrical charge.

When an atom emits an alpha particle its atomic number decreases by two units, since the alpha particle carries away two protons from the nucleus. Its mass number, however, goes down four units, since the mass number of the departing alpha particle is four (2 protons + 2 neutrons).

Emission of a beta particle, on the other hand, has no effect upon the mass number, since electrons have almost no weight. The atomic number, nevertheless, is actually increased, since the beta particle is formed by a neutron being transformed into a proton and an electron. The electron is shot out as a beta particle, while the proton remains in the nucleus, adding one to the number of protons and hence to the atomic number.

Gamma rays are generally emitted whenever a beta particle is ejected from the nucleus, and they sometimes also accompany alpha-particle emission. When someone rocks a boat on a quiet pond, waves spread out from it in all directions, and we may think of the gamma rays as the waves produced by an agitated nucleus when it has just thrown a particle overboard.

Let us consider an atom of radium as a specific example. The atomic number is 88, and the mass number 226. If we wait long enough, our atom of radium will emit an alpha particle. As just mentioned, this lowers the atomic number two units, and the mass number four units. Thus the new atom has an atomic number of 86 and a mass number of 222. It is an atom of the element radon, an entirely different substance in physical and chemical properties from radium. Its nucleus is still unstable, however, and in time another alpha particle is sent out, leaving an atom of atomic number 84 and mass number 218, an isotope of the element polonium.

But the end is not yet, for a third alpha particle is ejected, leaving an isotope of lead, atomic number 82, mass number 214. This nucleus is still unstable, but this time a beta particle is thrown out, and the atomic number rises to 83, the mass number remaining at 214. The result is an isotope of bismuth. After further successive emissions of alpha and beta particles, the atom ends up as a stable isotope of lead, atomic number 82, mass number 206.

Table I — A Radioactive Transformation Series

Particle emitted	Atomic number	Mass number	Isotope of
	88	226	Radium
alpha ← ↓	86	222	Radon
alpha ← ↓	84	218	Polonium
alpha ← ↓	82	214	Lead
beta ← ↓	83	214	Bismuth
beta ← ↓	84	214	Polonium
alpha ← ↓	82	210	Lead
beta ← ↓	83	210	Bismuth
beta ← ↓	84	210	Polonium
alpha ← ↓	82	206	Lead

We show in Table I the succession of transformations from radium to the end of the series, but to make the story complete we should have started with an isotope of uranium which by a similar series of particle emissions gives rise to radium itself. In addition to this so-called uranium series, two other similar series are known.

Our description of the uranium series shows that in any mass of radioactive material we actually have quite a variety of radioactive elements, some emitting alpha rays and some beta. Furthermore, any radioactive sample is not permanent, but continually diminishes in amount. Let us now answer some questions about the speed of radioactive disintegration.

Suppose we start with any weight of pure radium, say one gram. Then after about 1600 years, only half of it will remain, the other half having been converted into the lower members of the series. If now we put away the remaining half gram of radium and wait another 1600 years, half of it will disintegrate, and we will then have only one-quarter of a gram. Another 1600 years will see the remainder again cut in half, leaving us only one-eighth of a gram. After every 1600 years whatever amount of radium we had at the beginning of that period will have been cut in half, and you can see that although our sample is steadily decreasing in amount, it will take forever before absolutely all of it is gone.

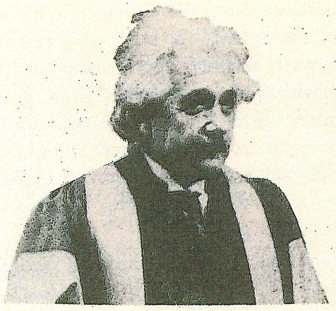
The 1600-year period in which the amount of any sample of radium decreases to one-half is called the half-life of radium. Other radioactive elements have widely different half-lives, ranging from billions of years down to a small fraction of a second. Table II gives, among other information, the half-lives of a number of typical radioactive substances, some belonging to the uranium series and some to the others.

Table II — Some Typical Radioactive Elements

Element	Atomic number	Mass number (atomic weight) of isotope	Emits	Half-life
uranium	92	238	alpha	4.45 billion years
radium	88	226	alpha	1600 years
polonium	84	210	alpha	140 days
actinium	89	227	beta	13.5 years
radon	86	222	alpha	3.8 days
proto-actinium	91	231	beta	32,000 years
polonium	84	216	alpha	0.14 second
thorium	90	232	alpha	18 billion years
thorium	90	230	alpha	83,000 years
thorium	90	231	beta	25 hours
thorium	90	234	beta	24.5 days

We have no way of predicting when any particular individual atom is going to explode, but we can make definite statements about the average behavior of the very large numbers of atoms composing the samples we study. We are in about the same position as a life insurance company with regard to the life spans of its policy holders. No one can predict how long any particular man will live, but if we have a group of 10,000 men of a certain age, we can be fairly sure how many of them will be alive a year from now. We don't know *which ones* will die, but we know *how many* will. In the same way, we do not know *which* atoms of a sample

of radium will disintegrate in the next 1600 years, but we are certain that half of them will.



Dr. Albert Einstein

We have already pointed out that the energy liberated in radioactive charges is many times greater than that involved in ordinary chemical reactions. What is the source of this tremendous energy?

As far back as 1905 Albert Einstein reached, by theoretical reasoning, the conclusion that mass and energy are equivalent. He suggested that it might be possible to find a process in which mass is converted into energy or energy into mass. His theory predicted that in such a process the amount of energy obtainable from a given amount of mass would be given by the very simple equation:

$$E=mc^2$$

in which E stands for the energy, m for the mass, and c for the speed of light.

When applied to any particular case, using the proper units, this equation gives amazing results. For example, if we could convert a dime (weighing 2.5 grams) completely into energy, the energy, according to Einstein's equation, would be sufficient to lift a weight of 15,700,000 tons to a height of one mile.

We now know that the energy observed to come from radioactive atoms is derived from the actual conversion of matter into energy. Einstein's equation shows that in order to account for the energies of the alpha, beta, and gamma rays, only a very small amount of mass need be used up, only a tiny fraction of a mass number unit, an insignificant percentage of the total.

In spite of the very great energy released, the natural radioactive elements are of no importance as general sources of energy, since those which are even relatively abundant disintegrate and release their energy with extreme slowness (remember the 1600-year half-life of radium), and no way is known of speeding up their decomposition.

In artificially produced atomic fission, on the contrary, we now have a source of even greater energy, and what is more, of energy which can be controlled. This matter is to be discussed in a later chapter.

CHAPTER IV

THE CYCLOTRON AND OTHER ATOM-SMASHING DEVICES

The high-speed alpha and beta particles shot out from the nuclei of radioactive atoms have been very valuable tools in the study of how atoms are put together. Recall from Chapter I that it was by bombarding nitrogen with alpha particles that Rutherford accomplished the first artificial transmutation of one atom into another. The discovery of the neutron and of artificial radioactivity were other consequences of the use of such particles in research.

Unfortunately, the radioactive elements, such as radium, which give us these particles are rare and expensive, and furthermore, they give us little choice as to type and energy of particle. We must take them as they come. It is small wonder, then, that physicists looked for artificial means of producing high-speed particles in larger amount and under better control.

The problem is two-fold: first, to produce the particles, and second, to speed them up so that they will have enough energy to be useful as bullets in bombarding atoms.

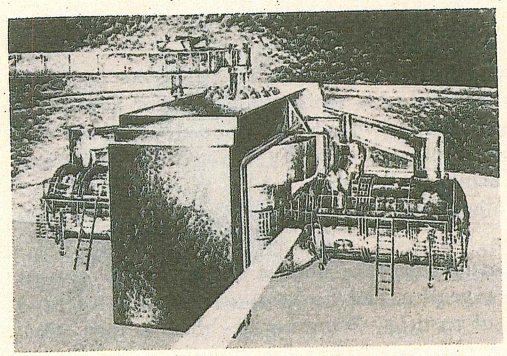
The first part of the problem is relatively easy. Recall that the atom of hydrogen consists of a nucleus, which is a single proton, around which revolves just one electron. To obtain protons, then, all we need to do is remove the electrons from hydrogen atoms, that is ionize them. The nucleus of the heavy hydrogen isotope, deuterium, contains in addition to the proton, also a neutron, and this nucleus is known as a deuteron. Thus to obtain deuterons, we simply ionize heavy hydrogen atoms. If we want alpha particles, we ionize helium atoms, since the alpha particle is a helium nucleus. The production of neutrons is another matter, which we will touch upon later.

To remove electrons from hydrogen, deuterium, or helium atoms so as to obtain their naked nuclei, we may shoot electrons at them. A heated metal filament, such as that in a radio vacuum tube, emits electrons, and if the filament is given a negative voltage with respect to its surrounding, the electrons are repelled from it with a considerable velocity.

If the space around the filament contains hydrogen at a low pressure, the speeding electrons, in colliding with hydrogen atoms, will frequently knock out of the atoms the electrons which they contain, leaving hydrogen ions, that is, protons in a free state. Under similar conditions, deuterium will yield deuterons, and helium, alpha particles. An electric arc may also be used to produce the ions.

Having formed the desired ions, we must set them in rapid motion by

giving them a large amount of energy. The most obvious method would be to place two metal plates in the apparatus and by means of a battery or generator, keep one of them charged positively and the other negatively. Then the positive ions will be repelled from the positive plate and attracted by the negative plate, and these forces will set them in motion.



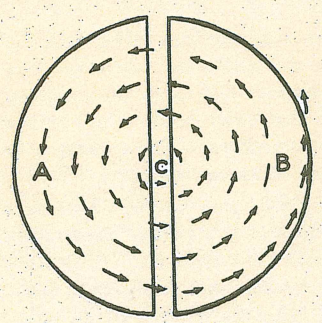
The difficulty here is that in order to give the ions high enough energies to make them useful in atom-smashing experiments, extremely high voltages are necessary, at least a million volts, and preferably much more. The production and control of such voltages are difficult, but several forms of apparatus have been built for that purpose.

The most famous of these is the Van de Graaf generator. In it, electricity is placed upon a rapidly moving paper belt by an electrical apparatus and carried by the belt to a large hollow metal ball supported on an insulator far above the floor. A potential difference of several million volts may be built up in this way between the ball and the ground. When this enormous voltage is applied to the two ends of an accelerating tube, ions produced at the positive end of the tube attain a very high speed by the time they reach the other end.

The difficulties and hazards of such extremely high voltages may be avoided if we can contrive to apply to our ions a succession of relatively small voltages. The most successful apparatus of this type is the cyclotron, developed about 1932 by Lawrence in California.

In the cyclotron the ions are made to move around a spiral path. The heart of the apparatus is a round, flat metal box, cut into two halves. These halves are called dees because of their shape like two capital letters "D" placed back to back. The dees are located flatwise between the poles of a very powerful electromagnet, weighing perhaps hundreds of tons. They are enclosed in an air-tight jacket from which the air can be removed by a vacuum pump.

In operation, positive ions are produced at the center of the dees (point C in the diagram). Dee A is made positive and dee B negative, 10,000 volts or more being used. The positive ions therefore move toward B. As they move, however, they are deflected by the magnetic field and forced into a semicircular path. Hence they quickly return to the gap between the dees. By this time, the electrical polarity of the two dees has been reversed, so that A is negative and B is positive. The ions are thus pulled across from dee B to dee A, and the pull increases their speed so that they move in a larger semi-circle than before. By the time they leave A, the polarity has again been reversed, and they are pulled back to B with another increase in speed. After an ion has been pulled back and forth from one dee to the other in this way several hundred times, its speed is as great as if it had fallen all at once through a single voltage difference of many million volts. Its path, furthermore, has gradually widened until it is at the edge of the dees, from where it may be led off for use in bombarding atoms.



The Dees of a Cyclotron

Protons, deuterons, and alpha particles may be given tremendous energies in the cyclotron. If we want neutrons, however, we must give them energy in some other way, for since the neutron carries no electrical charge, it is neither attracted nor repelled by the charged dees of the cyclotron.

Originally neutrons were obtained from atomic nuclei by bombardment with alpha particles from radioactive substances. The cyclotron now furnishes us with a convenient and powerful source of alpha particles, so in spite of our inability to speed up neutrons directly in a cyclotron, we can obtain them indirectly by allowing a beam of high-speed alpha particles (or other ions) from the cyclotron to fall upon a suitable element, generally the light metal beryllium. Neutrons in great abundance are then expelled from the beryllium nuclei.

Neutrons obtained indirectly from the cyclotron were of great importance in the studies which led up to the artificial release of atomic energy. The next chapter will explain why neutrons are of special interest as bullets for bombarding atoms.

For reasons connected with their very small weight, electrons cannot be speeded up to a useful extent in a cyclotron. The induction accelerator, or betatron, invented in 1940 by Kerst, makes up for this deficiency. By means of a rapidly changing magnetic field, it accelerates electrons to speeds greater than those of any natural beta particles. When high-speed electrons strike a solid metal target, they produce X-rays. Electrons from the betatron are capable of yielding X-rays which are "harder," that is, of shorter wave length, than those from ordinary X-ray tubes. It is likely also that these very energetic electrons will have interesting and perhaps useful effects upon atomic nuclei, but this line of research is just beginning to be opened up.

CHAPTER V

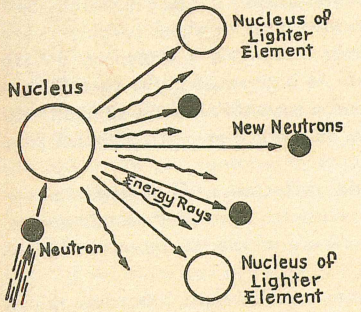
ARTIFICIAL ATOMIC ENERGY

In the chapter on atomic structure we saw that particles called neutrons are one of the building blocks from which the nuclei of atoms are constructed. They are similar in weight to protons, but carry no electrical charge. In the preceding chapter we learned that neutrons may be knocked out of the nuclei of certain light atoms such as beryllium. We come now to the part played by neutrons in the production of atomic energy.

Since neutrons have no electrical charge, they pass without hindrance among the electrons of atoms. Hence they have much greater penetrating power than the electrically charged alpha and beta particles. A neutron is stopped only when it comes so close to a nucleus that it scores a direct hit. In most cases, when a neutron collides with the nucleus of an atom, it is simply captured by the nucleus, and there is formed a nucleus which belongs to an isotope, often radioactive, of the old element.

There are a few atoms of a particular kind called fissionable isotopes, which do an especially interesting thing when hit by neutrons. These fissionable atoms are isotopes of the heavier elements, such as uranium, thorium, etc. The nuclei of these atoms are more complex in structure and contain more neutrons and protons than do atoms of the lighter elements. Thus their cohesive force is less and they are more easily split.

NEUTRON SPLITTING AN ATOM



When a neutron strikes the nucleus of a fissionable isotope, the entire nucleus splits into two pieces. The two fragments move away from each other with tremendous velocity and the energy rays which are simultaneously released appear ultimately as heat. The fragments collect an electron cloud around themselves and thus form new radioactive isotopes of other known chemical elements, such as barium. Apparently not every fission results in exactly the same fragments. It is this phenomenon of nuclear fission which constitutes the basis of atomic energy development.

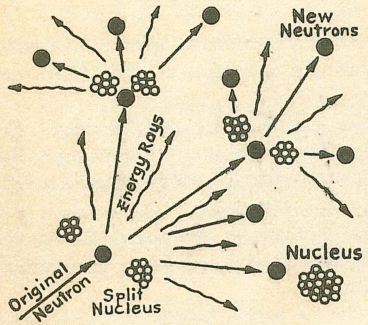
Nuclear fission, or the splitting of an atom by a neutron into two other elements, releases *artificial* or man-controlled atomic energy. The relatively slow natural disintegration of radioactive materials, such as uranium or radium, is a *natural* release of atomic energy.

If the high-speed neutron, which is sometimes referred to as an atomic bullet, can be slowed down in speed, the chances of its striking the nucleus of an atom are much greater. Several elements, among them carbon and deuterium (heavy hydrogen), have been found to exert a moderating or slowing effect on neutrons. This is the reason that graphite (carbon) is mixed with the element whose atoms are to be split.

The isotope of uranium of mass number or atomic weight 235 (U-235) is readily split by a slow-moving neutron. When this occurs additional neutrons are released, which in turn may split other atoms of uranium 235. At the same time, if atoms of uranium 238 are present with the atoms of 235 (which they always are) some of the neutrons released from the fission of the U-235 atoms are absorbed by atoms of U-238. This produces another isotope of uranium, called U-239. The U-239 atoms in turn emit a beta particle becoming neptunium, which in turn emits another beta particle becoming plutonium.

We see, therefore, that atoms of uranium 235 can be split by neutrons with the release of more neutrons, and at the same time, uranium 238 can be converted into plutonium.

CHAIN REACTION



This brings us to the so-called chain reaction. Neutrons produced by one nuclear fission fit and split other nuclei. But since atoms are mostly open spaces, a chain started in a small block of U-235 quickly dies out because most of the released neutrons escape from the block. The bigger the block, the smaller will be the percentage of escaping neutrons, and more will be left to split other nuclei. Thus, when a block of uranium 235 exceeds a certain size, the so-called critical size is passed and each 1000 nuclear fissions release sufficient neutrons to split more than 1000 additional nuclei. Under these conditions enough neutrons are produced to keep the reaction going and at the same time allow the extra neutrons to be absorbed by U-238 atoms, producing plutonium.

Plutonium is also a fissionable material and an atom of plutonium can be split by a neutron in the same way as can an atom of U-235. Thus, as the amount of plutonium formed in the reaction increases, the fission of atoms also increases and more neutrons are released.

This chain reaction is the basis of the atomic pile in which plutonium is produced and is also the basis of the atomic bomb. The difference is in the speed with which the reaction is allowed to proceed.

In Chapter III we learned that the source of the energy of radioactivity is a decrease in mass or weight of the atoms, in other words, the conversion

of matter into energy. The even greater energy released artificially in atomic fission likewise comes from the conversion of mass into energy.

When an atom of uranium is split by a neutron, two lighter atoms are formed along with several neutrons. The weights of these fragments all added together come out less than the weight of the original uranium atom plus that of the neutron which it captured. The difference is the mass which is converted into energy in accordance with Einstein's equation.

As nuclear fission is at present carried out, only about one-thousandth of the weight of a uranium atom is converted into energy when the atom is split. This is roughly one-fifth the weight of a neutron. If all, or even a large part of the mass of a sample of uranium could be changed into energy, our present atomic bombs would look pale in comparison.

CHAPTER VI

EXPERIMENTS IN RADIOACTIVITY

There are a number of very interesting experiments with radioactivity which you can perform with the materials and apparatus in your Chemcraft outfit.

The radioactive materials which are included are not harmful. The amount of radium is so small that the radioactivity is far below the point where it might be dangerous. Uranium is, of course, entirely safe to handle as its radioactivity is relatively low.

We learned in earlier chapters that radioactive materials, during the course of natural disintegration, give off three types of rays or emanations. These are known as Alpha, Beta and Gamma.

The alpha rays or particles are nuclei of helium atoms and are, therefore, particles consisting of two protons and two neutrons. Alpha particles have relatively large mass and hence do not penetrate as far as the other types of ray.

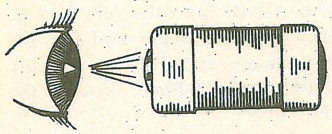
Beta rays are electrons, particles having very little weight or mass, but bearing negative charges of electricity.

Gamma rays are similar to X-rays but of shorter wave length and hence of greater penetrating power. It is the gamma rays which cause burns from exposure to intense radioactivity.

CHEMCRAFT ATOMIC ENERGY SPINTHARISCOPE

This special scientific instrument which is included in your outfit contains radium, one of the most valuable substances in the world. Radium costs about $25,000.00 per gram. When you look into your Atomic Energy Spinthariscope under proper conditions, you actually see the effect of the atomic disintegration of radium, the closest approach man has ever made to seeing an atom. This is accomplished by placing a small amount of radium compound close to a specially prepared screen. The alpha particles are expelled from the radium with a velocity of about 10,000 miles per second. As each particle strikes the screen, a flash of light is produced. These flashes, called scintillations, can easily be seen when observed through the powerful lens in the eye-piece of the Spinthariscope.

CHEMCRAFT SPINTHARISCOPE



Two rules must be carefully followed in order to successfully observe the flashes caused by the alpha particles.

(1) Your Atomic Energy Spinthariscope must be kept in a dark place for several hours before you use it. If any light is allowed to enter through the lens, the screen glows for some time. This of course makes it very difficult to see the alpha particle flashes.

(2) The instrument must be used in a dark place and you must remain in the dark for five or ten minutes before your eyes become adjusted. You will be unable to see the flashes until your eyes have become entirely adjusted to the dark. If necessary focus your Spinthariscope by moving the eye-piece in or out slightly.

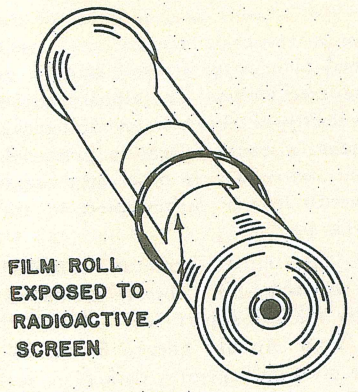
In a dark closet or at night when the lights have been turned out, hold the eye-piece of the Atomic Energy Spinthariscope to your eye and observe the continuous play of flashes on the screen. They will remind you of hundreds of bright stars twinkling in the dark sky. Most wonderful, however, will be the knowledge that you are seeing with your own eyes the effect of atomic disintegration and that the energy locked in the atoms of radium is so tremendous that this action will continue for thousands of years.

MAKING PICTURES BY ATOMIC ENERGY

The rays or emanations given off by radioactive materials affect photographic film or plates in very much the same manner as light rays. This fact opens a very interesting field for experiments from which much valuable information can be obtained.

There are several methods by which sensitive film or plates may be exposed to the emanations of radioactive substances. Great care should be taken against exposing the film or plates to light before they are developed, as the least exposure will probably spoil your experiment.

The films or plates used should be of medium sensitivity, that is, medium fast. Slow film or plates require longer exposure to the radioactive material, while fast film is difficult to handle because it is very sensitive to light.



The easiest method of making atomic energy pictures is to use an entire roll of film, removing only the metal foil wrapper, but not unrolling the film. Place your radioactive screen against the film roll and hold it in place with a small rubber band or by tying it with string. Now put the roll of film with the radioactive screen attached to it back into the film box or just wrap it up in paper and put the box or package in a drawer or any other dark place. Make note of the date and time.

Allow the radioactive screen to remain against the film roll for at least 24 hours. Heavier markings will be obtained if a longer time is allowed.

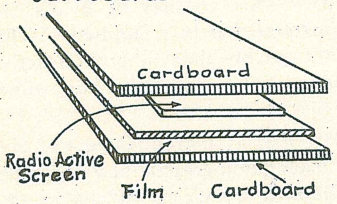
When you remove the radioactive screen from the film roll, either develop the film yourself or have it developed by a photographer, giving instructions that the film should not be cut into sections but is to be left in one piece. When you have the developed film, unroll it and hold the strip against a light background, such as a piece of white paper. If your film has not been fogged or light-struck, you will notice a very interesting pattern consisting of a series of black images of the radioactive screen. These will be regularly spaced and will gradually become lighter toward the end of the film that was on the inside of the roll. Study the pattern and visualize the penetration of the film by the rays emanating from the radioactive screen. It is possible that some of the rays may penetrate even beyond the center of the film roll, in which case you would have a secondary pattern in between the original pattern.

Since alpha, beta and gamma rays are all given off from uranium or its decomposition products, we might assume that each type of ray played some part in marking this film. However, alpha rays are not very penetrating, in fact they do not even pass through a heavy sheet of paper. Therefore, the markings on our film must be caused by the beta and gamma rays. Later on we will suggest experiments wherein certain other rays will be screened out so that the relative effect and, to some extent, the intensity of the different rays may be judged.

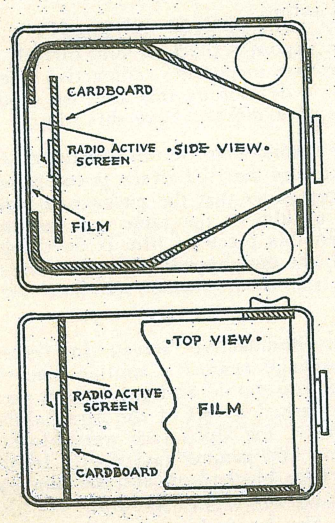
A number of methods of exposing film or plates to radioactive material will no doubt suggest themselves to the experimenter. We will set down a few ideas, however, which may be of practical value. Always be sure to place the radioactive material against the sensitive or coated side of the

film or plate. On roll film, the sensitive side is the inner side of the roll. If there is any doubts as to which is the coated side of a certain piece of film, snip off a small piece and take it into a light place. You will find the sensitive side is a dull cream color, while the non-sensitive is usually reddish and shiny. Possibly a simpler method is to place an edge of the film between your lips. The coated or sensitive side will stick to one lip, while the uncoated side will not stick to the other lip.

Radio Active Screen and Film placed between two Cardboards



Film may be cut into pieces in a dark room and the radioactive material placed against the film or against any desired type of screen which, in turn, is placed against the film. The entire package is then sandwiched between two pieces of cardboard, securely fastened and wrapped in dark paper. After the exposure time is over, open the package and remove the film, being very careful to keep it away from any light until the film has been developed. The above method is, of course, the best for handling plates.



In all cameras there is a space on the inside between the lens and the film or plate. A piece of cardboard can be cut to fit across this space and the radioactive material fastened to this cardboard by means of string or a rubber band. Use thin tissue paper or cheesecloth to wrap the radioactive material in if it is desired to obtain the alpha ray markings. The camera is then loaded with the film or plate in the usual way. Turn to the first exposure position and set the camera away for the desired length of time. Do not, of course, open the shutter. Additional exposures may be made by turning the film to the next exposure position.

Many radioactive materials will mark film or plates as described above. Uranium ores, such as carnotite or pitchblende, can be used, but due to the relatively low content of uranium, considerably longer exposure time must be allowed. The ore should be wrapped in a single thickness of thin

tissue paper or cheesecloth as the alpha rays do not penetrate through glass, plastics, metals, or even a sheet of fairly heavy paper.

Thorium, which is present in gas mantles, is radioactive. Gas mantles are still used in some types of kerosene, gasoline and acetylene lamps. If you can obtain a gas mantle, burn off the protective coating, then crush the mantle and use as described above for uranium ore.

Radium is much more radioactive than uranium. In the Chemcraft outfits, a Radioactive Screen is included. This consists of a disc to which a small amount of radium-bearing material has been applied. This radium disc will mark film very much more quickly than other radioactive materials mentioned because the emission of rays is many times more rapid.

We have learned that radioactive materials, as they disintegrate, give off three kinds of rays and that these rays mark photographic film or plates. In the following experiments, we will eliminate by means of shields first one type of ray and then another type, thus showing the relative amounts of the three rays emitted by radium and its radioactive decomposition products.

The alpha rays will not penetrate any dense materials, even heavy paper. Beta rays will penetrate thin layers of most materials, except lead. Gamma rays are the most penetrating, passing through almost all materials, including lead up to a thickness of several inches.

EXPERIMENT 1—First, suppose we expose a piece of film directly to our Radioactive Screen for 48 hours. Be sure to put the radioactive screen against the sensitive side of the film. Upon developing this film, we should have an intense black spot about the size of a nickel. Save this film.

EXPERIMENT 2—Next, wrap your Radioactive Screen carefully in a piece of heavy paper or cellophane. Make sure that the screen is completely inclosed and that the ends are securely creased so that the package is tight. Now place this package against a piece of film of the same type as used before and expose it for 48 hours. When this piece of film is developed, compare the mark with that obtained from the unwrapped radioactive screen. The black spot should be about the same size as the former spot but somewhat less dark.

EXPERIMENT 3—Now wrap your Radioactive Screen in several thicknesses of lead foil and place this package against a similar piece of film for 48 hours. When this is developed, you should find only a faint spot.

In the first experiment, the film was marked by all of the rays (alpha, beta and gamma). In the second experiment, the alpha rays were blocked off hence the film was marked by the beta and gamma rays only. In the third experiment, the alpha and beta rays were blocked off and the film was marked by the gamma rays alone.

The above series of experiments can also be performed with uranium ores, such as carnotite or pitchblende. Due to the much lower rate of radioactivity of these materials, an exposure time of several weeks must be allowed, instead of a day or two as with the radioactive screen.

Various salts of uranium and thorium, such as uranium oxide, uranyl acetate, uranyl nitrate, uranyl chloride, uranyl sulfate, thorium acetate, thorium nitrate, thorium chloride, thorium oxalate, thorium oxide and thorium sulfate, as well as metallic uranium and thorium, are sometimes available in laboratories. Small amounts of these salts of radioactive metals, of course, show radioactivity and if they are available, they can be used to mark film and to demonstrate the different types of rays emitted by radioactive materials. Uranium and thorium compounds containing these metals are much less radioactive than radium and its compounds and therefore require a longer time to mark film. Experiments with any of these materials should allow about 3 to 4 times as long an exposure as would be necessary with the radioactive screen.

In concluding this chapter, we wish to point out that the preceding experiments are merely suggestive, and we believe that the ingenuity of the experimenter will lead to many additional applications of the principles described.

CHAPTER VII

SOURCES OF RAW MATERIALS FOR ATOMIC ENERGY

The fissionable atoms of uranium 235 and of plutonium (made from uranium) have been used almost exclusively so far as the source of artificial atomic energy. It is known that certain isotopes of both thorium and protoactinium also undergo fission when bombarded by neutrons. Since uranium apparently has some advantages over thorium and protoactinium as a source of atomic energy, less consideration has been given to the latter two elements. Uranium, therefore, stands first as a source of atomic energy at the present time.

Although uranium is not considered to be a very rare material, relatively few deposits of commercial importance have been discovered up to the present time. Deposits of commercial interest from which uranium has been recovered in the past occur in the Belgian Congo, Canada, Czechoslovakia, Portugal and the United States. In addition, small amounts of uranium have been produced in England, Madagascar, and Russia.

The principal ores from which uranium is recovered are pitchblende and carnotite. Pitchblende (uraninite) is a complex oxide of uranium, containing also variable amounts of lead, iron, copper, bismuth and the rare elements, radium, thorium, yttrium, helium and argon. The percentage of uranium oxides in pitchblende varies from about 64 to 89 per cent. From

a physical point of view, pitchblende is an extraordinary mineral. From it helium was first obtained. The radioactive elements, uranium, radium, polonium and actinium were first discovered in pitchblende. The discovery and purification of uranium and radium from pitchblende led to many important discoveries in chemistry and physics.

Pitchblende has been found at various localities in the United States, mostly in very small quantities. There are small occurrences in Utah, California, Wyoming, Colorado, Nevada and Texas, also in the New England states. The principal deposits of pitchblende in the United States are in North Carolina, but even these are not large producers. The Great Bear Lake deposits in Canada were discovered in 1930, and immediately attracted world-wide attention. These are believed to be the richest deposits of pitchblende so far discovered.

Carnotite consists chiefly of potassium uranium vanadate. Carnotite is a soft, powdery mineral, generally a bright yellow color, but is sometimes discolored by iron oxides or organic matter. A sample of carnotite is included in your Chemcraft outfit.

Extensive deposits of carnotite occur in the plateau regions of Colorado and Utah. These deposits occur in connection with vanadium ore.

Radium is always associated in nature with uranium in the ratio of about one gram of radium to three tons of uranium. The principal interest in uranium ores before 1941 was due to the radium content. At this time radium sold for about $25,000.00 per gram, while uranium in the form of the impure oxide sold for only one or two dollars per pound.

Thorium is reported to be more abundant in nature than uranium. This element is rather widely distributed, occurring chiefly as thorium phosphate in monazite sand. Principal deposits of this ore occur in Brazil and India.

Of all these sources of atomic energy raw materials, pitchblende is at present by far the most important.

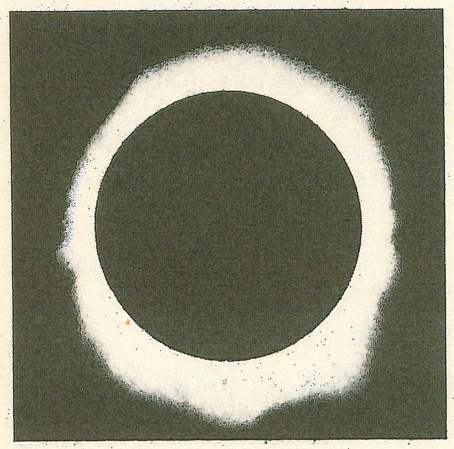
CHAPTER VIII

RADIOACTIVITY IN EVERYDAY LIFE

Certain natural phenomena heretofore clothed in the mysteries of nature have now become more clear to scientists due to our greater understanding of atomic energy. For one thing, radioactive atoms of heavy elements scattered through the depths of the earth are now believed to be playing an important part in preventing the earth from cooling off. Over millions

of years the gradual splitting of these atoms has been creating subterranean heat which has supplied the irresistible forces to push up new mountain ranges, cause volcanic eruptions, and change the shapes of continents.

It is quite probable that sometime in the distant past heavier elements existed in nature than we are able to find today. The fact that plutonium and neptunium disintegrate to form uranium, indicates strongly that these elements recently produced by man were once present in nature. In the future scientists will undoubtedly be able to explain many important facts about the formation and changes in our world, which are today clothed in secrecy.



Astronomers have applied the findings of the atom-smasher to the solution of one of the greatest problems about the sun and the stars. That is, what makes them shine and what supplies the apparently inexhaustible supply of heat. Our sun and the countless other suns throughout the universe could not produce their heat by burning in the usual sense. First, because calculations have shown that the sun would have burned away its substance long ago had its heat been produced by the usual oxidation. Second because of the temperatures on the sun, the atoms of oxygen and of oxidizable material are moving too fast to join together in the chemical union that constitutes burning. We now realize that the inexhaustible supply of heat and light emanating from our sun and other suns in the universe is the product of atomic disintegration.

We are all more or less familiar with the recent progress in the treatment of cancer with radium. It was discovered that cancerous tissue was more susceptible to the gamma rays given off by radium than was healthy tissue. Thus it has been possible in many cases to destroy or arrest the growth of cancerous tissue by radium treatment.

The extremely penetrating gamma rays from radium pass through several inches of steel. They may therefore be used for detecting flaws and cracks in heavy forgings and castings in much the same way that X-rays are used for locating breaks in bones and cavities in teeth. A small capsule of the radioactive preparation is fastened against one side of the metal object and a photographic film is placed against the other side. Since the rays pass

more readily through even tiny hidden cracks, the presence of such flaws is indicated on the film.

Another interesting application of radium with which we are more or less familar is radium luminous paint, which is used on markers and signs that must be seen in the dark, and also on the hands and dials of watches and various instruments. Radium luminous paint retains its brightness indefinitely, while other kinds of luminous paint are merely excited by exposure to light, but after a few minutes begin to lose their brilliance and gradually become dark until again excited by re-exposure to light. This property of radium luminous paint to shine continuously is based on the property of the radioactive rays to excite luminous calcium sulfide and cause it to glow in various degrees of brightness. Thus, if a small amount of radium is thoroughly mixed with luminous calcium sulfide and this mixture applied with a suitable vehicle to a surface in the form of a paint, a visible glow of light will result continuously.

Many other uses of atomic energy in our daily life will unfold in the future. Our knowledge in this field of science is in its infancy, and we may confidently expect great progress.

CHAPTER IX

THE DEVELOPMENT OF THE ATOMIC BOMB

In March 1939, only a few weeks after the discovery that uranium atoms could be split by neutrons, the possible military importance of fission was called to the attention of the government. In 1940, censorship was clamped on all atomic energy experiments, and in the summer of 1942, the entire project was consolidated under the name of "Manhattan Project." This project was probably the most important scientific and industrial achievement in the history of the world.

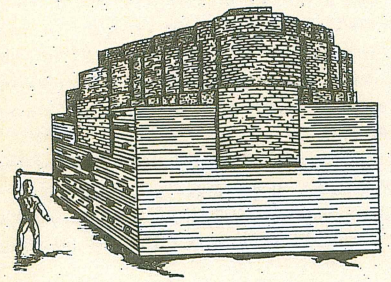
The practical application of atomic energy was at first pursued along two lines, both of which depended on uranium as the raw material.

One approach was to concentrate the fissionable U-235 isotope of uranium. This isotope occurs in all natural uranium, being .7% of the total, most of the balance being uranium 238. It had been worked out theoretically, that if a certain concentration of the U-235 isotope could be obtained in sufficient quantity to equal the critical size, spontaneous fission or an atomic explosion would result.

Two practical processes for separating U-235 from U-238 were developed and used on a large scale. One of these processes is known as the Gaseous Diffusion Process, based on the fact that a gaseous compound of U-235 will

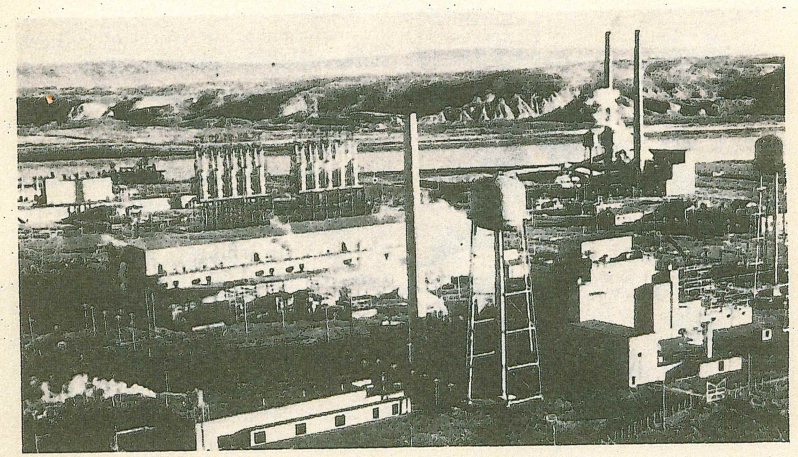
pass through a porous wall or barrier a little more readily than will the same compound of U-238. The other separation process is known as the Electro-Magnetic Process, wherein the natural mixture of U-235 and U-238 is vaporized and passed through an electric arc, where the atoms receive an electric charge, and then through an electric field where the atoms are given a high velocity. These atoms then pass into a magnetic field where the lighter atoms of U-235 curve more sharply than the heavy U-238 atoms, thus making possible the collection of the gaseous atoms in separate containers. Neither of these processes accomplished 100% separation, but these methods served to produce sufficiently high concentrations of uranium 235 for the purpose required.

THE FIRST ATOMIC PILE



The other approach to the practical production of a fissionable material was that of converting uranium 238 into plutonium, which as we have learned is fissionable like uranium 235.

In order to produce plutonium, it was necessary to construct a self-sustaining chain reaction pile. Such a pile was constructed at the University of Chicago in the fall of 1942. The pile was built on the lattice principle, that is, with lumps of metallic uranium separated by pieces of graphite.



Hanford Washington Plant

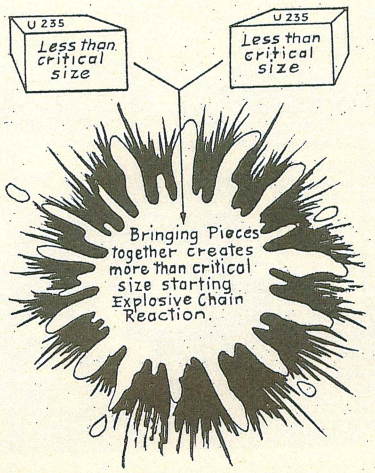
The purpose of the graphite being to slow the speed of the neutrons so that they could be more readily absorbed by the nuclei of the uranium 238 atoms.

At intervals in the pile, rods of cadmium or boron steel were inserted in slots. The purpose of these rods was to absorb neutrons should the reaction proceed too rapidly. These rods can be withdrawn or inserted as necessary.

The pile was shaped somewhat like a doorknob and was several times as high as a man. As the pile neared completion the critical size was reached, meaning that the neutrons released from atoms of U-235 in the uranium set up a chain reaction striking other atoms of U-235, releasing additional neutrons and continuing with a surplus of neutrons to convert some of the U-238 atoms into plutonium.

The successful operation of the experimental atomic energy pile resulted in the preparation of plans and the construction of larger piles for the production of plutonium in the necessary quantities. One of these piles was built at Oak Ridge, Tennessee, near Knoxville, and the others at Hanford, Washington, on the Columbia river.

With the availability of plutonium in sufficient quantities from the atomic piles and also with facilities for separating U-235 from U-238, two fissionable materials were at hand for the construction of the atomic bomb. This work was carried on in New Mexico, a short distance from Santa Fe. It is probable that bombs were made both from the concentrated U-235 and also from plutonium, the plutonium bomb being the more powerful.



An amount of fissionable material (either U-235 or plutonium) above the critical size will explode almost immediately. A stray neutron could start the chain reaction and there are many sources of neutrons within these fissionable materials, in addition to the possibility of neutrons from outside sources. Accordingly, the fissionable material must be kept in separate pieces, each of which is below the critical size. The essential problem in constructing the bomb was to arrange the separate pieces of fissionable material so that they could be brought together instantaneously at the time the bomb was to explode.

The explosion of the atomic bomb is therefore a chain reaction wherein a neutron strikes the nucleus of a fissionable atom, splitting it apart and in turn, releasing other neutrons to strike other atoms, and so forth. When the critical size referred to above is obtained, this chain reaction spreads so rapidly as to be instantaneous. The release in one fraction of a second of the atomic energy produced by the fission of the mass of material is beyond the force of any explosion heretofore created by man.

An atomic bomb produces great material damage by the terrific blast wave set up in the atmosphere. It is the shock of air pressure which breaks in walls and roofs of buildings and causes most of the damage to property. Intense heat and light are also produced from the explosion. This, of course, is the cause of uncontrolled fires which follow an atomic bomb blast.

The atomic bomb, completely unlike other explosives, emits an enormous number of gamma rays at the moment it explodes. When an atom undergoes fission, not only are fragments and neutrons released, but in addition, much high-energy radiation. Therefore, at the time the bomb explodes, people in its vicinity are exposed to what amounts to a severe over-dose of X-rays, which cause disintegration of the tissues. The fragments split from atoms of the bomb which are left over after the bomb has exploded are radioactive. They continue to emit radioactive rays for several months. This means that the vicinity of the explosion of an atomic bomb is unsafe for human habitation for some time after a bomb has exploded.

CHAPTER X

THE PRINCIPLE OF THE HYDROGEN BOMB

At this point it might be well for the reader to refresh his memory on the structure of the atom as described in Chapter II, especially the portion pertaining to isotopes. We know quite accurately the proportionate weights of practically all the known atoms and their isotopes. All of these weights are very close to whole numbers, as would be expected from the fact that all atoms are composed of protons and neutrons weighing almost exactly one unit apiece, and of electrons weighing almost nothing. But although the atomic weights are close to whole numbers, they are not exactly whole numbers, and the differences are important.

If we take the difference between the atomic weight of an isotope and the nearest whole number and divide this difference by the whole number, we obtain the so-called "packing-fraction" of the isotope. The packing-fraction is related to the mass which is converted into energy when the nucleus is built up out of its constituent particles.

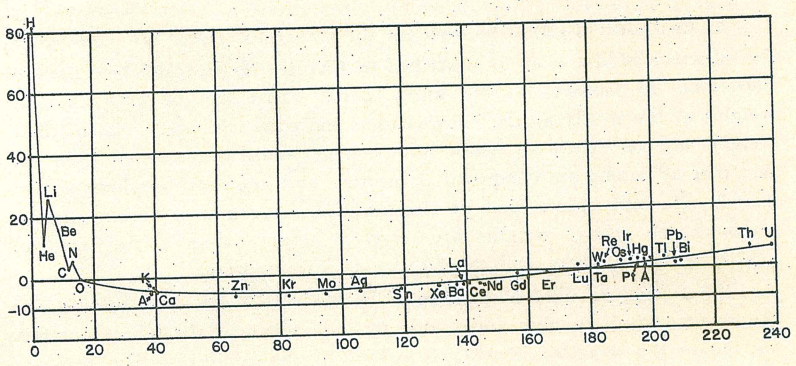
Let us arrange the elements in order of atomic weight and plot vertically on a graph the packing-fraction of each. The result is the packing-fraction curve shown in the drawing. The very light nuclei have large packing-fractions, and the very heavy nuclei also show an increase in packing-fractions as compared with those in the middle. The line of zero packing-fraction passes through oxygen simply because oxygen is the standard of the system of atomic weights.

The packing-fraction curve shows that if an atom of uranium splits apart to form two atoms near the middle of the curve, energy is released. This is the source of the energy of the atomic bomb. If, however, we could cause several atoms of hydrogen to combine so as to form an atom with a lower packing-fraction, still greater energy (per unit weight) would be available. This is the source of the energy of the sun and stars, and it is this energy which it is proposed to tap in the hydrogen bomb.

Two forces are at work inside of an atom. First, an electrical repulsion which makes the positively charged protons repel each other as similar electric charges usually do. Second, the so-called binding force which causes protons and neutrons to attract each other at very short distances equivalent, say, to the diameter of an atomic nucleus. At greater distances, the electrical repulsion force becomes the main force at work and the binding force falls off rapidly as the distance between a proton and neutron increases.

Suppose you could take a small assortment of neutrons and protons and squeeze them closely enough together for the binding force to act. They would snap into an atom, releasing energy as they did so.

PACKING–FRACTION CURVE



To make a very large atom, you would have to shove the protons and neutrons in forcibly because the diameter of the big atom is getting close to the limit of attraction of the binding force and the electrical repulsion between protons is almost as strong as the binding attraction. Such an atom would of course be very unstable, which is one of the reasons that the very heavy atoms are the ones which exhibit natural radioactivity.

Certain of the heaviest atoms, such as plutonium and the uranium isotope of mass 235, may be caused to split apart by neutrons striking their nuclei. Two middle-weight nuclei are produced and several new neutrons are released. The energy liberated depends upon the difference in the packing-fractions of the original atom and of the new ones. Under suitable conditions the neutrons released by splitting atoms of uranium or plutonium produce still more neutrons and set up in this way a self-sustaining chain reaction.

Now let us go to the other end of the packing-fraction curve where, as we have learned, we obtain energy by putting light atoms together instead of disintegrating heavy atoms. It is quite evident that the amount of energy contained by combining a gram of hydrogen atoms is vastly in excess of the energy obtained by disintegrating a gram of uranium-235 atoms.

It may sound very simple to compress hydrogen atoms and neutrons so that they would snap together forming helium and release enormous amounts of atomic energy. The difficulties, however, are tremendous. Ordinary hydrogen fuses into helium only in a complicated series of steps much too slow to be of interest on the earth, though it is in progress on a gigantic scale in the sun.

It is believed, nevertheless, that either, or both, of the two rarer isotopes of hydrogen, deuterium and tritium, could be made to form helium by subjecting them to the extremely high temperature reached in the explosion of an ordinary atomic bomb. The deuterium nucleus contains a proton and one neutron, and that of tritium a proton and two neutrons. Although these two isotopes have all the chemical properties of hydrogen, their nuclear properties are entirely different.

Deuterium is readily available since 1 part in 5000 of natural hydrogen is deuterium and it can be separated by electrolysis of water. Tritium, however, would have to be manufactured in uranium piles from an isotope of lithium, Li6. The lithium isotope, Li6, when bombarded with neutrons, would form helium and tritium.

The temperature in an atomic bomb explosion is measured in millions of degrees and is probably high enough to touch off the fusion reaction of deuterium and tritium. Of course in a very small fraction of a second, the bomb begins to expand and to cool. A major problem is, therefore, how to keep the deuterium and tritium atoms from being scattered by the explosion before they have a chance to participate in it. In order to cram enough of these atoms into a small space they would have to be liquefied

by cooling nearly to the absolute zero, and they would have to be kept cold until the bomb was exploded.

The power of the hydrogen bomb is apparently limited only by the quantity of hydrogen isotopes which can be made available within the bomb under conditions suitable to their fusion into helium.

CHAPTER XI

THE FUTURE OF ATOMIC ENERGY

The entire field of atomic energy is in its exciting infancy. The tremedous power locked in the atom was, out of war-time necessity, first released for purposes of destruction. But today the great scientists in the leading laboratories of the world are bending every effort to adapt this great force to constructive peacetime uses.

Because these more laudable efforts on the part of scientists to harness this great power are in the early stages of development, it is impossible to foresee the results that will be arrived at in the future. We do know, however, that the constructive forces of atomic energy are being applied experimentally in the following fields of science.

POWER PRODUCTION

Many suggestions have been offered for the use of atomic power. Among these are power and heat supply for cities and large industrial plants, power plants for ships, and power plants in remote places to operate mines or to recover other natural resources. From the point of view of the public welfare, the introduction of such a new source of power is a clear gain. If it could lessen the cost of heat and power to our cities, it would be a stimulus to every industry. If it could give cheaper power where industry and agriculture need it but cannot now get it, it would extend our economic frontiers. These are possibilities that future research and development may make realities.

TAGGED ATOMS

The hundreds of radioactive isotopes now available as by-products from atomic energy piles will prove to be very important in our future lives. Chemically indistinguishable from the ordinary form of the element, these isotopes serve as tagged atoms when mixed with stable atoms of the same element. The atoms of the radioactive isotopes mix and move with the stable atoms, and they serve to identify and mark atoms of this particular element wherever they may travel. Identification of the tagged atoms is possible because they are radioactive and emit alpha, beta and gamma rays, which may be detected by a sensitive screen, a film or by a Geiger counter.

BIOLOGY AND MEDICINE

Present developments indicate that the greatest promise of atomic energy in biology and medicine is in the use of radioactive tracers or tagged atoms and of intense radiations. For the first time, the opportunity is now at hand to attack thoroughly an almost infinite number of problems which are of concern to the fundamental mechanisms of life processes. This is accomplished by the introduction of radioactive isotopes of various elements into the various functioning parts of animal and plant organisms, and by this means, to follow and analyze the functioning of their organs.

The application of radiation therapy should be widened by the great variety of radioactive products made available by the development of atomic energy. While there is no evidence at the time that the rays from these artificial radioactive materials will supersede X-rays, at the same time, we may expect great progress in this general field.

ENGINEERING

In the design and testing of machinery and in many other phases of engineering, radioactive tracers or tagged atoms and other radioactive products have promising applications. Radioactive tracers could be applied to problems of wear and lubrication of moving parts. They could be used for tracing flows of liquid or gaseous materials. They might be employed for reducing static electricity, particularly in explosives plants and in the printing industry. Radioactive sources of gamma rays could be used as substitutes for industrial X-ray machinery. Such a source of rays being small in size could be used in places where an X-ray machine would not be practical.

METALLURGY

In metallurgy the use of tagged atoms again offers many possible applications. The diffusion of an element into itself or into alloys of which it is a component could be followed if the diffusing atoms were radioactive. Identification of traces of impurities, which often affect the properties of metals and alloys, could be made by adding a radioactive form of the suspected impurity and photographing the resulting mixture by microradiographic methods.

CHEMISTRY AND CHEMICAL ENGINEERING

Radioactive materials, it is believed, will come to play an important part in industrial chemistry. Now that these materials are less expensive and in much wider variety of chemical forms, more uses will be found for them. The availability of a variety of radioactive tracer elements will offer possibilities for solving questions heretofore difficult to answer, since radioactive atoms of a given element will travel with their non-radioactive brothers, giving a positive means of identifying this element in molecular and in colloidal forms. The radiations from atom-splitting chain reactions are capable of inducing chemical reactions which otherwise could not be carried to completion.

PHYSICS

The new possibilities for producing high-energy gamma rays as well as fast and slow neutron beams, and alpha and beta rays in intensities vastly greater than ever before, will aid in the solution of the basic unsolved problems of modern science; the structure of matter and the nature of radiation. It may also be possible to study the properties of matter under immensely high temperature conditions, somewhat resembling those in the stars.

THE GREAT UNKNOWN

Fifty years ago it was evident that X-rays were useful for seeing through objects, such as the human body, which are opaque to ordinary light. It could not be predicted that X-rays would become a powerful weapon in the fight against cancer, or that the study of X-rays would lead to the discovery of radioactivity and stimulate the development of radio, radar, and a host of electronic devices. Such unforseen developments are the result of every great discovery. It would, therefore, be surprising if the most important consequences of the release of atomic energy are not in directions as yet unpredictable.

The Story of

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SAFE, INTERESTING EXPERIMENTS

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$E = mc^2$

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GAMMA

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TO THE MYSTERIES AND MIRACLES
OF AN ATOMIC WORLD

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CHEMICAL MAGIC

Mystifying Magical Demonstrations

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CHEMICAL MAGIC

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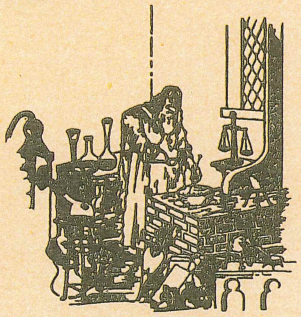
INTRODUCTION

Centuries ago, before Chemistry came to be known as a science, members of the Zoroastrian order of Magi mystified the masses with their seemingly supernatural powers.

During the period from 300 B. C. to 200 A. D. under the inspiration of Hermes Tristmegistus widely separated groups of individuals experimented with many weird concoctions, believing that they could change such base metals as iron, lead and copper into precious gold.

Since heating causes many chemical changes to take place, the most popular experiment of those ancient men was to heat or burn various substances and mixtures. Based on these experiments, the Arabic word "al" means 'the' and the Greek word "chemia" meaning 'to heat' or burn were combined, producing the word "alchemy." In those days people were very superstitious, and so the alchemists who had learned to bring about such wonderful changes in materials came to be regarded as wizards or magicians.



AN ALCHEMIST

Eventually, the alchemists gave up their fruitless attempts to produce gold and endeavored to devise a mixture known as "the elixir of life." This mythical substance they believed would prolong their lives so that they could live to be as old as Methuselah.

Gradually as time progressed and as people gained more knowledge the experiments and the ideas of the old alchemists passed into the discard. However, some of the things which the alchemists attempted to do in a very crude and unscientific manner have become the basis for many modern theories. Numerous achievements of present day scientists seem like miracles of magic until we understand the scientific principles upon which they depend. The intriguing feats of modern day magicians appear weird and supernatural until you learn how they are performed. Every one enjoys being mystified, and the skillful tricks originated by such magicians as Blackstone, Thurston and Houdini never cease to hold audiences enthralled.

In the pages that followed you will find tricks in chemical magic which will provide hours of entertainment. By using simple chemical reactions you can amaze your friends by producing beautiful color changes, thick clouds of smoke without fire, diabolical odors, invisible inks and scores of other tricks which are equally mysterious. Complete directions are given for presenting a Chemcraft Magic Show including suggestions for stage settings, costumes, suitable "patter" and publicity.

SECTION I

EXPERIMENT 1—How to Change Water to Sherry Wine.

Pour 4 or 5 drops of tincture of iodine into a dry glass tumbler and allow it to become dry. The tumbler is now prepared to perform the magic change and you can proceed as follows: "Friends, I am going to change plain water to sherry right before your eyes." At this point pour some water into your prepared glass, mumble a few magic words over it, wave your hand, saying, "Presto, Change-o", and the water will immediately be changed to an amber-colored sherry wine.

To heighten the illusion you can prepare several glasses with iodine beforehand and instead of making only one glass make several glasses of sherry for your friends. Naturally, you cannot drink this wine.

EXPERIMENT 2—How to Change Sherry to Water.

To a tumbler ½ full of water add 2 measures of Sodium Thiosulfate (No. 16). Pour this solution into ½ tumbler of the sherry made in the preceding trick, and immediately the glass will be filled with a colorless liquid which you can call water.

This effect is quite startling to an audience if they do not understand the decolorizing of sodium thiosulfate on iodine solution.

EXPERIMENT 3—Pouring Red and White into the Same Tumbler.

Dissolve 3 measures of Sodium Carbonate (No. 4) in a glass full of water, and add 2 or 3 drops of Phenolphthalein Solution (No. 22). In another glass dissolve 2 measures of Sodium Bisulfate (No. 7) in 3 or 4 drops of water. Now pour the red solution in the first glass into the second. At first the red liquid will turn colorless, but as you add more it will change to red again.

EXPERIMENT 4—Pouring Grenadine and Water into the Same Tumbler.

Put 4 measures of Sodium Bisulfate (No. 7) in a glass full of water and add 2 or 3 drops of Phenolphthalein Solution (No. 22). In another glass put one measure of Sodium Carbonate (No. 4) and add just enough water to dissolve the solid.

Now pour from the first glass into the second. At first a red color will result, but as you add more of the solution from the first glass the red color will mysteriously disappear and apparently only water is in the glass.

MYSTERIOUS LIQUIDS

EXPERIMENT 5—How to Change Water to Burgundy Wine and Burgundy to Water.

Fill two tumblers half full of water and add to each 1 or 2 drops of Phenolphthalein Solution (No. 22). To the first one add also ⅓ measure of Sodium Carbonate (No. 4).

Now one glass contains a red liquid (wine) and the second glass apparently plain water. To the red liquid add 1 measure of Tartaric Acid (No. 14) or a little vinegar, and to the second or colorless liquid add ⅓ measure of Sodium Carbonate (No. 4). Now the colors have become reversed, the red liquid being in the second glass and the colorless liquid in the first.

To change them back again, add to the first glass 3 measures of Sodium Carbonate (No. 4) and the second 2 measures of Tartaric Acid (No. 14) or some vinegar.

EXPERIMENT 6—Changing Water to Blood and Blood to Solid.

Put one inch of Sodium Silicate Solution (No. 19) in a test tube and add water until the tube is ⅓ full. In another test tube put 2 or 3 drops of Phenolphthalein Solution (No. 22) and fill ⅓ full of water. In another test tube ⅓ full of water put 5 measures of Aluminum Sulfate (No. 12) and shake the tube until dissolved.

Now hold the three tubes in your hand and pour the contents all at once into a clean glass. The three colorless liquids coming together immediately form a red blood-like liquid which in a few minutes will turn into a red solid mass.

EXPENIMENT 7—A Magic Pitcher of Grenadine.

This experiment is especially good to entertain your friends:

Obtain a good-sized pitcher; a glass pitcher will not do for this experiment, as you need one which is not transparent. Place the pitcher on a tray or table, fill

it with water and arrange 5 glasses around it. In one glass put 1 measure of Sodium Carbonate (No. 4), in another 3 drops of Phenolphthalein Solution (No. 22) and in the third put 1 measure of Sodium Bisulfate (No. 7). The remaining glasses are used to make the experiment seem more complicated to the audience and need not be used unless you wish. Add a few drops of water to each glass to dissolve the chemicals. Be sure to remember which glass contains the sodium bisulfate.

Now you can show the pitcher to the audience and let them see that it contains only water. Fill each of the glasses with water from the pitcher and then pour the contents of each glass back into the pitcher, except the sodium bisulfate solution. Speak some magic words over the pitcher and then pour the contents into the empty glasses. The liquid will be wine .

Now empty all of the glasses into the pitcher again, this time including the sodium bisulfate solution, then after you have cast your magic spell over the pitcher pour the contents into the glasses for a third time, when it will be clear water.

EXPERIMENT 8—A Magic Pitcher of Bluing.

Arrange a pitcher and glasses again as in a preceding experiment. Put 1 measure of Sodium Ferrocyanide (No. 6) in one glass, 1 measure of Ferric Ammonium Sulfate (No. 21) in another, and in still another 6 measures of Sodium Carbonate (No. 4).

Fill the pitcher with water and pour into the glasses. Empty all of the glasses back into the pitcher except the one containing the sodium carbonate solution.

Now you have a pitcher of bluing. Pour this into the empty glasses and empty all of them back into the pitcher again, this time including the sodium carbonate solution.

Fill the glasses again from the pitcher and you will apparently have only water.

EXPERIMENT 9—A Magic Pitcher of Wine and Bluing.

Arrange a pitcher and 4 glasses as in the preceding experiment. In the first glass put 1 measure of Sodium Carbonate (No. 4), in the second glass a few drops of Phenolphthalein Solution (No. 22), in the third glass put 2 measures of Ferric Ammonium Sulfate (No. 21) and in the fourth glass put 2 measures of Sodium Ferrocyanide (No. 6).

Now put enough water in the pitcher to fill the glasses about half full and pour the water from the pitcher into the glasses.

Pour the contents of the two glasses back into the pitcher, say a few magic words, and then pour the liquid into the glasses again, and the result will be wine. Now pour the contents of the other two glasses back into the pitcher, say some different magic words, pour the liquid into the glasses again and it will be bluing.

EXPERIMENT 10— Magic Blue Writing Paper.

Put ½ measure of Sodium Ferrocyanide (No. 6) and ½ measure of Ferric Ammonium Sulfate (No. 21) on a smooth white piece of paper. Mix the chemicals together. Lay the paper on a smooth surface, and rub the mixture over the entire piece of paper. Now shake off the extra powder and write on the paper with water, using a clean pen. The water writes just like beautiful blue ink.

Some people will not believe that you can write with plain water, but just fix up some sheets of paper in this way and let them try it. To further mystify them have a few untreated sheets to try beforehand.

EXPERIMENT 11—Magic Black Writing Paper.

Lay a sheet of ordinary writing paper on a smooth flat surface, put ½ measure of Tannic Acid (No. 15) and ½ measure of Ferric Ammonium Sulfate (No. 21) on it and mix them together. Using a small piece of paper rub this mixture over the entire sheet. Your hands and also the paper must be perfectly dry, as the slightest trace of moisture will cause black smudges. When the mixture has been rubbed over the paper, shake off the remaining powder and write on the prepared surface with water. The written characters will be black just as if you had used black ink.

EXPERIMENT 12—Magic Writing Fluids.

Set 3 glasses in a row and number them 1, 2 and 3. Put 4 teaspoonfuls of water in each glass.

In the first glass dissolve 1 measure of Tannic Acid (No. 15).

In the second glass dissolve 1 measure of Ferric Ammonium Sulfate (No. 21).

In the third glass dissolve 1 measure of Sodium Ferrocyanide (No. 6).

Number three pieces of writing paper 1, 2 and 3. Using a clean pen, write on each piece of paper with the solution in the glass of the same number, thus paper

number 1 would be written on with the solution in glass number 1. Be sure to wipe your pen carefully each time you use it for a fresh solution.

When the writing on the papers has dried, wet paper number 1 with solution number 2, using a small soft brush or sponge to apply the solution (be sure to wash the brush or sponge thoroughly when changing it from one solution to another). Wet part of paper number 2 with solution number 1, and part of it with solution number 3, wet paper number 3 with solution number 2.

The many colors in which the writing appears makes this a very mystifying experiment to show your friends.

EXPERIMENT 13—How to Prepare Magic Black Blotting Paper.

Prepare a solution of 1 measure of Tannic Acid (No. 15) in a test tube ¼ full of water. Heat the solution for a few minutes to dissolve the solid, and then write with it on a sheet of smooth white paper, using a clean pen.

While the writing is drying prepare a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ½ full of water, and moisten a small piece of clean blotting paper and the writing will appear black just as if written with ink.

EXPERIMENT 14—A Magic Blue Blotter.

Write on a sheet of paper with a solution of 1 measure of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. When the writing is dry, blot it with a piece of blotting paper moistened with ferric ammonium sulfate solution as in the preceding experiment, and bright blue writing will appear on the blank paper.

EXPERIMENT 15—Making Black Writing Appear with a Magic Pen.

Prepare a solution of 1 measure of Tannic Acid (No. 15) in a test tube ¼ full of water. Heat this solution for a minute or two to dissolve the solid. Now using a clean pen, write on a sheet of paper with this solution as ink.

Prepare another solution of 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Clean your pen and follow as nearly as you can the writing which you did with the tannic acid. As you follow the tannic acid lines with ferric ammonium sulfate solution they will become black. It is very interesting to see how closely you can follow the other lines, and you can work out very nice puzzles for your friends with this experiment by writing on various parts of a sheet of paper with the tannic acid solution, then make some ferric ammonium sulfate solution, tell them to find what you have written by drawing lines around the paper until they come to your writing.

EXPERIMENT 16—Making Blue Writing Appear with a Magic Pen.

Dissolve 1 measure of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. Use this solution as ink and write on a sheet of white paper. Now prepare a solution of 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water and trace out the writing with a pen using the ferric ammonium sulfate solution as ink. This time the lines will appear blue when you go over them.

EXPERIMENT 17—Writing that Disappears by Magic.

Prepare a solution of ½ measure of Ferric Ammonium Sulfate (No. 21) and ½ measure of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. Shake the solution thoroughly, and then, using a clean pen, write with the blue ink on a sheet of white paper.

Prepare another solution of 2 measures of Sodium Carbonate (No. 4) in a test tube ¼ full of water. Dip your pen in this solution and follow along the blue writing. The writing will disappear as you go over it with the sodium carbonate solution, making a very interesting trick.

EXPERIMENT 18—Salt Writing.

Dissolve 6 measures of common table salt (sodium chloride) in a test tube ½ full of water. Write with this solution on white paper using a clean pen. When the writing is dry, scratch over it with a soft lead pencil and the words, which you have written will show up plainly in dark lines.

EXPERIMENT 19—Blue Printing by Magic.

Put in a test tube 1 measure of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21). Fill the test tube ⅓ full of water and shake the solution thoroughly. Now paint a piece of white paper with the blue

solution. When the paper is entirely colored allow it to dry, and prepare a solution of 2 measures of Sodium Carbonate (No. 4) in a test tube ¼ full of water.

Now draw on the blue paper with the sodium carbonate solution and your lines will appear white, making an effect very similar to a blue print.

EXPERIMENT 20—Developing Water Writing with Ink.

Obtain a clean pen and write on a hard finish paper using ordinary tap water. Make your letters rather bold and wide and allow this water writing to dry thoroughly and then paint it with any kind of a thin ink. Next, quickly take up the ink with a blotter and note how the characters which you have written with water stand out.

This is sometimes the way secret messages are written and developed, especially in espionage work.

SYMPATHETIC INKS

Writing done with sympathetic ink is invisible until it has been developed by warming or treating it with some chemical reagent. Such inks have frequently been used by spies in time of war for conveying military secrets through enemy territory. Although their practical value is limited, these inks form the basis for a number of very fascinating magic tricks.

EXPERIMENT 21—How to Make Yellow Sympathetic Ink.

Put 3 measures of Ammonium Chloride (No. 9) in a test tube. Add ½ inch of water and shake it until the solid is dissolved. Now write with this solution and when the writing is dry heat the paper. The lines will appear yellow when the paper is just slightly scorched.

EXPERIMENT 22—Rust-Colored Sympathetic Ink.

Dissolve 4 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Write with this solution.

To make the writing appear dissolve 4 measures of Sodium Carbonate (No. 4) in a test tube ¼ full of water and carefully spread this solution over the writing, using a soft brush or tuft of cotton. The writing will appear a red-brown color just as if written with iron rust.

EXPERIMENT 23—Sympathetic Ink from Lemon Juice.

Write with the juice of a lemon, using a clean pen. The writing will be invisible until it is brought out by heating the paper.

EXPERIMENT 24—Grape Fruit Juice as Sympathetic Ink.

Write with the juice of grape fruit, making the lines heavy. Upon heating, the writing will appear and if it is scorched it will turn brown.

EXPERIMENT 25—Using Vinegar as Sympathetic Ink.

Vinegar can be used as a sympathetic ink. Upon heating the writing will appear. It does not fade out on cooling.

EXPERIMENT 26—Sympathetic Ink from Milk.

Write with milk, being careful to make your lines rather heavy. Upon heating the paper, writing will appear.

EXPERIMENT 27—Iron Sympathetic Ink.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Writing done with this solution will appear brown when the paper is heated.

EXPERIMENT 28—Making Sympathetic Ink from an Onion.

Obtain some of the juice of an onion and write with it on white paper using a clean pen. When heated the written characters turn to a brown color and do not fade away. This is a very stable sympathetic ink.

EXPERIMENT 29—How to Make White Writing on Black Paper.

Dissolve 6 measures of Sodium Bisulfate (No. 7) in a test tube ⅓ full of water and dip a small piece of white paper into the solution until it is thoroughly wet. When the paper is dry write on it with a solution of 2 measures of Sodium Carbonate (No. 4) dissolved in a test tube ¼ full of water.

To bring out the writing heat the paper until it is scorched, being very careful not to set it afire. The paper will turn black and the writing will appear in white lines.

EXPERIMENT 30—Blue Writing Appears and Disappears.

Dissolve 2 measures of Cobalt Chloride (No. 27) in a test tube ¼ full of water, and write with this solution on a sheet of white paper. When the writing is dry, warm the paper gently, being careful not to scorch it, and the letters will appear in blue. To make them disappear, breathe on the writing for a minute or two, or hold it in the steam rising from boiling water.

CHEMICAL WITCHCRAFT

EXPERIMENT 31—Making a Chemical Clock.

Prepare in several glasses a series of solutions of Sodium Thiosulfate (No. 16) of decreasing strength; for example, place in the first glass 10 measures of sodium thiosulfate, in the second 6 measures, in the third 3 measures, and in the fourth 1 measure. Fill each glass half full of water and stir until the solid is dissolved.

Prepare a solution of 6 measures of Sodium Bisulfate (No. 7) in a test tube of water, and add an equal portion of this to each glass. A milky precipitate of colloidal sulfur will appear after a few minutes, but it will form in the different glasses in different lengths of time according to the concentration of the thiosulfate solution. A little experimenting makes it possible to cause the appearance of the precipitate within a specified time.

EXPERIMENT 32—A Chemical Smoke Screen.

Put 3 or 4 measures of Ammonium Chloride (No. 9) on a spoon and heat it over a flame. In a few moments it will begin to give off thick clouds of smoke which will continue until the substance is entirely volatilized.

EXPERIMENT 33—Magic Ice.

Pour ½ inch of Sodium Silicate Solution (No. 19) in a test tube and add 3 or 4 drops of water. Shake to mix the liquids.

In another test tube put 4 measures of Sodium Bisulfate (No. 7) and fill ¼ full of water. Shake it until the solid is dissolved.

Pour the liquid in the first tube into the second. The mixture will gradually "freeze" into a transparent solid, becoming hard first at the surface and gradually solidifying through the entire mass. In a minute of two the test tube can be inverted as there will be no liquid to spill.

EXPERIMENT 34—How to Make Bones Elastic.

Obtain a small chicken bone such as the wish bone. Clean it thoroughly and cover it with strong vinegar for 24 hours. At the end of this time remove the bone from the vinegar and it will be so elastic that you will be able to tie knots in it without breaking it.

EXPERIMENT 35—How to Stretch an Egg.

Obtain an egg and put it in a glass full of vinegar and allow it to stand for about 24 hours. At the end of this time feel the egg with your fingers and if it is soft and elastic, remove it from the vinegar and examine it carefully. If it is not soft at the end of this time pour out the vinegar and replace it with fresh vinegar for another 24 hours.

When the egg has become elastic it may be put in a bottle through which it originally would not pass without breaking. An elastic egg will be quite a novelty to show your friends, but be careful not to break it.

EXPERIMENT 36—Disappointing Matches.

Using a small brush, paint around the bases of the heads of matches with Sodium Silicate Solution (No. 19). Apply a thick coating of the sodium silicate solution and paint the stick for a distance of about one-fourth of an inch from the head. Allow the matches to dry and then try to light one. It will strike all right,

but the flame will go out almost immediately if you have put on enough sodium silicate solution.

EXPERIMENT 37—Blowing a Colorless Liquid to a Magic White.

Dissolve 2 spoonfuls of sugar in a glass of water and add 2 measures of Calcium Oxide (No. 20). Stir for three or four minutes to dissolve nearly all of the calcium oxide. Now let the solution stand for a few minutes until the upper portion is perfectly clear.

Pour out some of the clear upper portion into a clean glass, filling it about ¼ full. Blow through the clear liquid for several minutes, using a glass tube. The colorless liquid will gradually become white.

Here is a good chance to play a joke on some of your friends who do not know about chemical tricks. Work the above experiment for them and ask them if they would like to try it, but instead of giving them the solution with sugar and calcium oxide prepare a solution of sugar and Sodium Carbonate (No. 4). They could blow into this all day without getting the slightest change.

EXPERIMENT 38—Blowing a Red Color to a Magic White.

Prepare a solution of sugar and Calcium Oxide (No. 20) as in the preceding experiment and pour some of the clear liquid into a clean glass, filling it about ¼ full. Now add to this clear liquid one or two drops of Phenolphthalein Solution (No. 22). It will turn red.

Blow through the red liquid, using a glass tube, for four or five minutes. It will begin to change color almost immediately, passing into various shades of red and pink until it finally becomes pure white.

EXPERIMENT 39—Blowing Red to Colorless.

Place ¼ measure of Sodium Carbonate (No. 4) in a glass ¼ full of water, add one or two drops of Phenolphthalein Solution (No. 22) and the liquid will become red.

Using a glass tube, blow through this red liquid for a few minutes. It will suddenly become colorless and will look just like ordinary water.

EXPERIMENT 40—A Fading Color.

Fill a milk bottle or clear glass jar about ½ full of water and dissolve in the water ½ measure of Sodium Carbonate (No. 4). Now add 1 or 2 drops Phenolphthalein Solutuion (No. 22) and the water will be colored pink.

Place in a glass a teaspoonful of baking soda or washing soda and add to this a little vinegar. An effervescence will take place and carbon dioxide gas will be given off. Hold the glass over the milk bottle containing the pink solution and tip it as though you were pouring from it into the jar or bottle, but do not allow any of the liquid contents of the glass to run out. The gas which is given off, however, will pour into the bottle or jar, as it is heavier than air. Continue pouring this gas into the jar for four or five minutes and add a little more vinegar to the baking soda or washing soda if necessary.

Now cover the mouth of the jar with a saucer or lid and let it stand for a few hours. Look at it now and then after awhile you will find that the color is disappearing in spots. Before long the color will be gone without any apparent cause.

EXPERIMENT 41—A Mysterious Blue Trick.

Put about one quart of water in a clean pan and heat to boiling on the stove. While the water is heating, put a lump of starch about as big as the end of your thumb in a small pan and mix it up to a smooth paste with a few drops of water. Add a little boiling water to the starch paste and then pour the paste into the pan of boiling water. Allow to boil for a minute or two. Remove from the stove and cool.

When your starch solution is cool pour it into a milk bottle or large glass pitcher or jar. Now carefully add 2 or 3 drops of tincture of iodine, to the cold starch mixture, and watch the color form. You can now shake or stir the liquid and add a few drops more of the iodine until the desired shade of blue is obtained. By pouring the blue liquid back into the pan and heating to boiling, the blue disappears, and you have a colorless liquid again.

By putting 9 or 10 drops of tincture of iodine into a milk bottle full of water and adding a little starch mixture you can produce this same blue color. It is easy to make a trick out of this and produce a blue color by pouring two colorless liquids together.

EXPERIMENT 42—A Synthetic Cow.

Fill a glass ¼ full of water and add 8 measures of Sodium Thiosulfate (No. 16). Stir until the solid is completely dissolved.

In another glass ¼ full of water dissolve 8 measures of Sodium Bisulfate (No. 7).

Pour the solutuion of sodium bisulfate into the glass containing the solution of sodium thiosulfate and let it stand for about 5 minutes. A precipitate will form, giving the solution the appearance of milk.

This experiment is especially good for exhibition because the solution may be mixed behind the scenes and the glass of apparently clear water can then be set before the audience and after awhile it will gradually turn milky. The milk mysteriously appears without anyone touching or going near the glass.

EXPERIMENT 43—Conjuring Red from Blue.

Fill two glasses each half full of water. In the first glass put 2 or 3 drops of Phenolphthalein Solution (No. 22). In the second ⅓ measure of Ferric Ammonium Sulfate (No. 21), and ⅓ measure of Sodium Ferrocyanide (No. 6); stir the contents of each glass slightly.

In the third put a spoonful of Sodium Silicate Solution (No. 19) and 2 spoonfuls of water; stir slighlty to mix.

Now take the first glass containing the colorless liquid in one hand and the second glass containing the blue liquid in the other hand; pour them both at once into the sodium silicate solution. A glass full of bright red liquid will result.

COLORED WIZARDRY

EXPERIMENT 44—The Magic Handkerchief.

Prepare a solution of 2 measures of Cobalt Chloride (No. 27) in a test tube ¾ full of water. Obtain a small handkerchief and put it in the test tube so that it soaks up the cobalt chloride solution.

Remove and dry the handkerchief on a heater or over a stove and you will find that it is blue when thoroughly dry and warm. Here's where you have some fun.

Show the handkerchief to someone while it is blue, and then crush it between your hands and blow through it for a few minutes. It will become colorless. You can repeat this as often as you like, turning the handkerchief blue by again warming.

EXPERIMENT 45—A Color Thermometer.

Place 1 measure of Cobalt Chloride (No. 27) in a test tube and add 2 drops of water. A pink solution will be formed. Heat the tube gently over a flame and the solution will turn deep blue in color; on cooling the color again changes to pink.

This change can be repeated as often as you wish, although it may be necessary to add another drop of water occasionally to replace that which is lost by evaporation.

EXPERIMENT 46—Mystifying Opposite Changes.

Take a small strip of Congo Red Paper (No. 32) and a small strip of Blue Litmus Paper (No. 18), hold the strips side by side and drop them into a glass half full of water in which 2 measures of Sodium Bisulfate (No. 7) have been dissolved. The blue litmus paper wil' turn red and the congo red paper will turn dark blue.

Now prepare a solution of 1 measure of Sodium Carbonate (No. 4) in half a glass of water and drop the papers into this solution. The dark blue congo paper will turn back to red, while the red litmus paper will turn blue.

EXPERIMENT 47—A Contrary Solution.

Put 2 measures of Sodium Carbonate (No. 4) in a clean glass and fill it about ¼ full of water, stir well to dissolve the solid. Add 1 or 2 drops of Phenolphthalein Solution (No. 22). The liquid will be colored red.

Now dip a piece of red Litmus Paper (No. 18) into the solution just made. It will turn blue. A piece of blue litmus paper dipped into the red solution will remain blue.

This trick is very simple from a chemical standpoint but to any one not familiar with indicators it looks very strange to set a piece of red paper turned bright blue by a red solution.

EXPERIMENT 48—The Mysterious Picture.

To a test tube ¼ full of water add 3 measures of Cobalt Chloride (No. 27) and shake until it is entirely dissolved. Using a small, soft brush, paint a picture on a sheet of paper using this solution. Dry the paper over an electric heater or a stove, being very careful not to scorch it. When the paper is warm and dry, the picture will appear in blue. To make it disappear again, breathe on it for a minute or two, or hold it in the steam from a boiling tea-kettle. You can make the picture appear and disappear as often as you like.

EXPERIMENT 49—Green Violets.

Place 2 measures of Ammonium Chloride (No. 9) and 2 measures of Calcium Oxide (No. 20) in a dry test tube and heat them over a flame. Hold a blue violet at the mouth of the tube as you heat it, and the gas which comes off will turn the flower green.

If you try other kinds of flowers you may obtain some weird and unusual results.

DIABOLICAL ODORS

EXPERIMENT 50—How to Make a Disagreeable Odor.

Place 4 measures of Sulfur (No. 1) in a dry test tube and add a piece of candle about ⅛ inch long. Heat the test tube over a flame, and after a minute notice the disagreeable odor of the evolved gas. The gas is hydrogen sulfide, which has an odor resembling that encountered when rotten eggs are broken.

EXPERIMENT 51—A Magic Odor that will Revive Fainting Persons.

Place 2 measures of Ammonium Chloride (No. 9) and 2 measures of Calcium Oxide (No. 20) in a test tube. Heat this mixture over a flame for a few moments and smell the gas which is given off. This gas has the property of reviving people who have fainted.

EXPERIMENT 52—How to Produce Odor of Violets.

Mix 3 or 4 drops of castor oil with 1 measure of Sodium Carbonate (No. 4) and heat this in a test tube over your alcohol lamp flame. After a short time you will notice a white cloudy fog rising in the test tube. Remove the tube from the flame, let it cool, and cautiously smell the vapors. You should be able to recognize the scent of violets.

CHEMICAL SORCERY

EXPERIMENT 53—The Mysterious Thread of Ashes.

Fill a test tube ¼ full of water and add sodium chloride (common table salt) until no more will dissolve when the test tube is shaken. Obtain a piece of white cotton thread two or three feet long and push it down into the test tube containing the salt solution, when thoroughly wet, remove the thread and hang it up to dry.

Now tie a very small nail or button or some other small object on one end of the thread and tie the other end of the thread to a hook or projection so that it will hang clear of the wall and way from draperies or anything that might catch on fire. Light the thread toward the lower end and it will burn up, but the ash remaining will still support the object in the air.

EXPERIMENT 54—A Magic Bleeding Wound.

Prepare a solution of 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water. Fill another test tube half full of water and add about 3 drops of Phenolphthalein Solution (No. 22).

Now moisten your hand or your arm with a few drops of sodium carbonate solution. When this has dried apply a small quantity of phenolphthalein solution to the same spot and the red color which is formed has the appearance of a serious wound. This experiment, of course, can be worked in a number of different ways. You can tie up your finger with a rag that has been moistened with the sodium carbonate solution and then apply a few drops of phenolphthalein solution to the rag.

To heal the wound, apply a few drops of a solution made by dissolving 1 measure of Sodium Bisulfate (No. 7) in a test tube ¼ full of water.

Be sure to wash all traces of the chemicals from your hands and arms when finished with this experiment.

EXPERIMENT 55—How to Develop an Invisible Picture.

Draw a line across the center of a sheet of writing paper and on one side of the line draw a picture using Phenolphthalein Solution (No. 22) as ink. Paint the paper on the other side of the line with Sodium Silicate Solution (No. 19). Fold the paper at the line and let the invisible picture touch the sodium silicate solution. Open the paper and the picture will appear in red lines.

EXPERIMENT 56—An Obedient Egg.

Obtain two one-quart fruit jars, and fill one of them ¾ full of water and the other with a concentrated salt solution, made by dissolving as much salt as possible in 1½ pints of hot water.

Now place an egg in the jar of water, and it will sink. Place it in the salt solution, and it will float. By remembering which jar contains the salt solution and which the water, you can make the egg obey you and either float or sink as you command.

EXPERIMENT 57—Water That Will Not Spill.

Obtain a piece of cardboard or stiff paper large enough to cover a tumbler. Fill a tumbler brimful of water and slide the piece of cardboard over it. Now hold the card in position and turn the tumbler upside down. You will find that the cardboard no longer needs to be held, for the atmosphere presses upon every square inch of it with a force of 14.7 pounds, and this is sufficient to hold it tightly against the mouth of the glass so as to prevent the water from falling out.

EXPERIMENT 58—The Mysterious Mothball.

Fill a small olive or pickle bottle ¾ full of vinegar, and add 7 or 8 measures of Sodium Carbonate (No. 4). Now drop in a mothball, and the bubbles which form on it will carry it to the surface, where it will turn around, and then sink again to accumulate another group of bubbles. This action will be repeated again and again.

SECTION II

The following additional tricks in chemical magic can be performed with Chemcraft Chemistry Outfits No. 3½, No. 5 and No. 10.

MAGIC CHANGES

EXPERIMENT 59—Pouring ink and Milk from the Same Pitcher.

In one glass place 1 measure of Tannic Acid (No. 15) and in a second glass place 2 measures of Strontium Chloride (No. 10). Add about a teaspoonful of water to each glass and stir until the chemicals are dissolved.

Now in a small pitcher or other vessel dissolve 5 measures of Ferric Ammonium Sulfate (No. 21) in about a glass of water.

When ready to exhibit pour half of this liquid into each of the first two glasses. The liquid in the first one will become intensely black like ink, and that in the second will become milky white.

EXPERIMENT 60—Pouring Red, White, and Blue from the Same Glass.

Fill 3 glasses ⅔ full of water. In the first dissolve 1 measure of Sodium Salicylate (No. 26). In the second dissolve 3 measures of Strontium Chloride (No. 10) and in the third dissolve ½ measure of Sodium Ferrocyanide (No. 6).

In another glass ⅓ full of water dissolve 5 measures of Ferric Ammonium Sulfate (No. 21). Pour some of this solution into each of the other 3 glasses and stir.

The liquid in the first glass will turn red, in the second white, and in the third blue.

EXPERIMENT 61—Pouring Red, Yellow, and Black from the Same Vessel.

Fill 3 glasses each half full of water and arrange them in a row. In the first glass put 1 measure ofTannic Acid (No. 15), in the second 5 measures of Sodium Bisulfate (No. 5) and in the third 1 measure of Sodium Salicylate (No. 26). Stir each glass until the solid is dissolved.

In a fourth glass full of water dissolve 5 measures of Ferric Ammonium Sulfate (No. 21). Stir until the solid is completely dissolved. Pour from the fourth glass into each of the other three. In the first black will be formed, in the second yellow, and in the third red. This makes a very pretty series of colors.

EXPERIMENT 62—Pouring Blue, Gold, and Red from the Same Vessel.

Fill 3 glasses each half full of water and dissolve in the first glass ½ measure of Sodium Ferrocyanide (No. 6), in the second glass 5 measures of Sodium Bisulfite (No. 5) and in the third 1 measure of Sodium Salicylate (No. 26).

In the fourth glass full of water .dissolve 5 measures of Ferric Ammonium Sulfate (No. 21).

Now pour from the fourth glass into each of the other three. In the first a blue color results. In the second a golden yellow and in the third a red.

THE TRANSMUTATION OF METALS

EXPERIMENT 63—Changing Steel to Copper.

Put 1 test tube of water into a clean glass and add 8 measures of Sodium Bissufate (No. 7) and 2 measures of Azurite (No. 39). Stir well to dissolve the solids. You now have a glass containing a magic blue liquid into which you may dip any piece of bright steel for a moment and have it covered with a bright red coating of copper. A knife blade makes an excellent article to use in this experiment. It should be rinsed in cold water and dried with a soft cloth after being removed from the solution. For best results the steel article to be plated should be bright and shiny.

EXPERIMENT 64—Changing a Dime to a Penny.

Dissolve 2 measures of Azurite (No. 39) and 5 measures of Sodium Bisulfate (No. 7) in a test tube half full of water. Pour this into a clean glass and place a bright silver coin and a small piece of iron in the solution, a small nail will do very well. Be sure that the iron touches the coin. In a few moments the coin will be covered with a red coating of copper.

By means of this experiment you can easily change your friends' dimes into pennies.

EXPERIMENT 65—Changing Iron Filings to Bronze Particles.

In a test tube ½ full of water put 1 measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7). Shake this mixture thoroughly until the solids are entirely dissolved and you have a clear blue solution in the test tube.

Next add 2 measures of Powdered Iron Metal (No. 36) and again shake the test tube for a few minutes. Now let the iron metal settle to the bottom of the test tube and carefully pour off the liquid. Examine the iron metal and you will find that it is coated with a deposit of copper and looks like particles of bronze.

MYSTERIOUS LIQUIDS

EXPERIMENT 66—How Water can be Changed to Magic Milk.

Fill 2 glasses half full of water and dissolve in one 2 measures of Manganese Sulfate (No. 24). In another dissolve 2 measures of Sodium Carbonate (No. 4).

Now pour the contents of one glass into the other and a thick white precipitate will be formed, making it seem as though the glass were full of milk.

EXPERIMENT 67—Changing Water to Wine and Wine to Milk.

Put 3 measures of Sodium Carbonate (No. 4) in a glass and fill the glass ¾ full of water, stirring to dissolve the solid. In another glass put 2 drops of Phenolphthalein Solution (No. 22), an din a third glass put 3 measures of Strontium Chloride (No. 10) and 3 measures of Sodium Bisulfite (No. 5). Pour the solution in the first glass into the second, and the result will be wine. Pour the wine into the third glass, stir vigorously for a few seconds, and it will change to milk.

EXPERIMENT 68—A Magic Pitcher of Milk and Ink.

Obtain a pitcher and four glasses. In the first glass place 3 measures of Mangenese Sulfate (No. 24), in the second glass 3 measures of Strontium Chloride (No. 10), in the third glass 2 measures of Tannic Acid (No. 15), and in the fourth glass 2 measures of Ferric Ammounim Sulfate (No. 21). Add a few drops of water to each glass, and pour four glasses of water into the pitcher.

Pour from the pitcher into each glass, and then pour the contents of the first two glasses back into the pitcher. Fill them again from the pitcher, and this time they will appear filled with milk. Now pour the second pair of glasses into the pitcher and when you pour again from the pitcher, the liquid will be ink.

MAGIC PAPERS AND INKS

EXPERIMENT 69—Magic Red Writing Paper.

Mix on a sheet of white paper ½ measure of Sodium Salicylate (No. 26) and ½ measure of Ferric Ammonium Sulfate (No. 21). Rub this mixture over the surface of the paper, using a small sheet of paper for rubbing. Write on the prepared paper with plain water, and the result will be beautiful red writing just as if you used red ink instead of water.

EXPERIMENT 70—Colorless Ink Writes in Three Color.

Make up three solutions as follows:

First, 1 measure of Sodium Ferrocyanide (No. 6) in a test tube half full of water.

Second, 1 measure of Sodium Salicylate (No. 26) in half a test tube of water.

Third, 1 measure of Tannic Acid (No. 15) in half a test tube of water.

Using a soft brush, paint these sloutions in streaks on the same piece of paper. Be careful to have each streak just touch the one next to it.

When the sheet of paper is thoroughly dry, write on it with a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of

water. The writing will be a different color on each streak, giving a many-colored writing on the same sheet of paper. Many interesting effects can be obtained by combining the streaks in various ways, or painting the streaks diagonally across the paper, or painting the solutions on in circles.

EXPERIMENT 71—Magic Writing in Different Colors.

Make up three solutions as follows:

First, 1 measure of Sodium Ferrocyanide (No. 6) in a test tube half full of water.

Second, 1 measure of Sodium Salicylate (No. 26) in half a test tube of water.

Third, 1 measure of Tannic Acid (No. 15) in half a test tube of water.

Using a soft brush, coat a piece of white writing paper with solution number 1, a second with solution number 2, and a third piece with solution number 3. Be sure to clean your brush well each time you use it for a fresh solution.

Set the three pieces of treated paper aside to dry and prepare a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in half a test tube of water.

When the sheets of paper are thoroughly dry, write on each one with the solution of ferric ammounium sulfate, using a clean pen. The same liquid writes a different color on each sheet of paper.

EXPERIMENT 72—Blue Writing with Iodine.

Put ½ measure of Ferric Ammonium Sulfate (No. 21) and ½ measure of Sodium Bisulfate (No. 7) into a test tube ¼ full of water and add a few drops of Sodium Iodide Solution (No. 23). Shake the tube well, and the solution will become brown because it contains free iodine.

Using a perfectly clean pen write on a starched color or cuff with this iodine solution. At first the writing will be blue, but in time it may turn brown. The solution will also write blue on paper which has been coated, or sized, with starch.

EXPERIMENT 73—Red Writing Appears When Blotted.

Write on a piece of paper with a solution of 1 measure of Sodium Salicylate (No. 26) in a test tube ¼ full of water. When the writing has dried, blot it with a piece of blotting paper moistened with ferric ammounium sulfate solution as before, and this time bright red writing will appear on the blank paper.

EXPERIMENT 74—Making Red Writing Appear with a Magic Pen.

Dissolve 1 measure of Sodium Salicylate (No. 26) in a test tube ¼ full of water and write with this solution on a piece of white paper. Prepare some ferric ammonium sulfate solution by dissolving 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Follow the writing with a pen dipped in ferric ammonium sulfate solution and the lines will appear red.

EXPERIMENT 75—A Magic Pen and Many-colored Writing.

Prepare in separate test tubes the following solutions: 1 measure of Tannic Acid (No. 15) in a test tube ¼ full of water; 1 measure of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water; 1 measure of Sodium Salicylate (No. 26) in a test tube ¼ full of water. Write on a clean sheet of paper with first one and then the other of these solutions making various designs, and having the writing done with one solution cross the writing done with the other solutions.

Next prepare a solution of 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Dip your pen in this solution and follow along the writing which you did on the paper. The writing will appear in various colors and you will find it a very interesting and fascinating experiment.

SYMPATHETIC INKS

EXPERIMENT 76—How to Make Light Red Sympathetic Ink.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. Write on a white paper with the solution.

Place in another test tube 1 measure of Azurite (No. 39), 3 measures of Sodium Bisulfate (No. 7) and ¼ test tube of water. Shake the tube gently until the solution stops foaming and becomes clear. Now spread this solution over the writing, using a soft brush or a small tuft of cotton, and it will appear in clear red characters.

EXPERIMENT 77—Dark Red Sympathetic Ink.

Dissolve 1 measure of Sodium Salicylate (No. 26) in a test tube ¼ full of water. Write with this solution.

In another test tube dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) and 3 measures of Sodium Bisulfate (No. 7). Shake the test tube gently until the solution stops foaming and becomes clear. Now spread this solution over the writing, using a soft brush or a tuft of cotton and clear red characters will appear.

EXPERIMENT 78—How to Make Dark Olive Green Sympathetic Ink.

Place in a test tube 1 measure of Azurite (No. 39), 3 measures of Sodium Bisulfate (No. 7) and 1 measure of Ammonium Chloride (No. 9). Fill the tube ¼ full of water and heat to dissolve the solids.

Write with this solution, using a clean pen. When the writing is dry, heat the paper over a flame and the invisible writing will appear a dark olive green.

EXPERIMENT 79—How to Make Blue Sympathetic Ink.

Dissolve 1 measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Write with this solution and when the lines have dried, prepare a solution of 1 measure of Sodium Carbonate (No. 4) in a test tube containing ½ inch of water. Paint the sodium carbonate solution over the writing, using a brush and the lines will appear a blue color.

CHEMICAL WITCHCRAFT

EXPERIMENT 80—Making a Miniature Snowstorm.

Dissolve 2 measures of Strontium Chloride (No. 10) in a test tube nearly full of water.

In a second test tube put 2 measures of Sodium Carbonate (No. 4) and add ½ inch of water. Heat until the sodium carbonate is dissolved.

Carefully pour the contents of the second test tube into the first. A thick white precipitate will form at the top of the tube and as the particles slowly settle to the bottom the effect is very much like a miniature snow storm.

Sodium carbonate and strontium chloride react forming sodium chloride and strontium carbonate. The latter is not soluble in water and hence separates out of the solution in solid particles.

EXPERIMENT 81—Chemical Color Chase.

Arrange 5 dry glasses in a row. Put in the first ½ measure of Ferric Ammonium Sulfate (No. 21), in the second ½ measure of Sodium Salicylate (No. 26), in the third ½ measure of Sodium Ferrocyanide (No. 6), in the fourth a spoonfull of Sodium Silicate Solution (No. 19), and in the fifth 2 or 3 drops of Phenolphthalein Solution (No. 22).

Fill the first glass with water and stir to dissolve the solid. Pour the contents of the first glass into the second, the second into the third, and so on. Stirring each time a change is produced will help bring out the color.

EXPERIMENT 82—Color Clots.

Put a spoonful of Sodium Silicate Solution (No. 19) in a glass and add 2 spoonfuls of water. Mix well.

Fill a test tube half full of water and add 3 measures of Strontium Chloride (No. 10). Shake the tube thoroughly. Pour a few drops of this solution from the test tube into the glass containing the sodium silicate solution. Beautiful white clots of precipitate will form.

Add 1 or 2 drops of Phenolphthalein Solution (No. 22) to the remaining strontium chloride solution. Pour a few more drops from the test tube into the glass. This time red clots of precipitate will form among the white ones.

EXPERIMENT 83—Streamers from a Magic Rainbow.

Fill a large glass pitcher, milk bottle or similar vessel of clear glass with water. Allow it to stand for ten minutes where it will not be shaken so that the water will become perfectly quiet.

Now drop 1 or 2 measures of Mixed Dyes (No. 30) on the surface of the water. Watch closely the beautiful color effects which follow.

MAGIC FLAGS

EXPERIMENT 84—"Old Glory" Appears by Magic.

Draw a flag on a sheet of white paper in light pencil lines, laying out the

blue field and stripes. Now make up the following solutions:

First, 1 measure of Sodium Ferrocyanide (No. 6) in a test tube half full of water.

Second, 1 measure of Sodium Salicylate (No. 26) in a test tube half full of water.

Third, 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water.

Paint the part of the flag which is to be blue with the first solution (sodium ferrocyanide). Paint the stripes which are to be red with the second solution (sodium salicylate).

Allow the work to dry thoroughly and then paint over the whole flag with the solution of ferric ammonium sulfate, using a soft brush or a small wad of cotton. The flag will come out in beautiful red, white and blue pattern of "Old Glory."

EXPERIMENT 85—The "Tri-Color" of France by Chemical Magic.

Draw a flag in light pencil lines and lay out three wide strips running parrallel to the staff or the short way across the flag. Paint the first of these stripes with sodium ferrocyanide solution prepared as in the preceding experiment. Leave the second strip blank and paint the third with sodium salicylate prepared as in the preceding experiment.

Paint over the whole flag with ferric ammonium sulfate solution as in the preceding experiment. The result will be the blue, white and red flag of France.

EXPERIMENT 86—The Flag of Holland.

Draw another flag in lead pencil, making three stripes, running the length of the flag. Paint the top stripe with sodium salicylate solution, leaving the middle stripe blank, and paint the bottom stripe with sodium ferrocyanide solution. When the painting is dry go over the entire flag with ferric ammonium sulfate solution. Prepare all the solutions as in the preceding experiment.

This time the flag of Holland results with three stripes, the first red, the second white and the third blue.

EXPERIMENT 87—Changing Magic Rainbow Streamers into a Floating Garden.

Dissolve 2 measures of Sodium Carbonate (No. 4) in a glass full of water. Allow the glass to stand until the water becomes perfectly quiet and then add 1 or 2 measures of Mixed Dyes (No. 30). Wait until the streamers of color from the dyes have reached the bottom of the glass, then add 2 measures of Aluminum Sulfate (No. 12) without stirring. As the alumnium sulfate dissolves it forms a precipitate with the sodium carbonate. This precipitate absorbs a large amount of the color from the streamers and carries it to the surface of the water. The changing colors and the motion in this experiment makes it very beautiful and interesting to watch.

SECTION III

The following additional tricks, in chemical magic can be performed with Chemcraft Chemistry Outfits No. 5 and No. 10.

MAGIC CHANGES

EXPERIMENT 88—Pouring 10 Different Colors from the Same Pitcher.

Arrange 10 dry glasses in a row. In the first put 1 measure of Ferric Ammonium Sulfate (No. 21) and 1 measure of Tannic Acid No. 15).

In the second put 3 measures of Strontium Chloride (No. 10) and 3 measures of Sodium Carbonate (No. 4).

In the third put 1 measure of Sodium Salicylate (No. 26) and 1 measure of Ferric Ammonium Sulfate (No. 21).

In the fourth put 3 measures of Cobalt Chloride (No. 27) and 2 measures of Sodium Carbonate (No. 4).

In the fifth put 1 measure of Calcium Oxide (No. 20) and 1 or 2 drops of Phenolphthalein Solution (No. 22).

In the sixth put 3 measures of Sodium Ferrocyanide (No. 6) and 3 measures of Nickel Ammonium Sulfate (No. 13).

In the seventh put 1 measure of Ferric Ammonium Sulfate (No. 21) and 4 measures of Sodium Carbonate (No. 4).

In the eighth put ⅓ measure of Sodium Ferrocyanide (No. 6) and ⅓ measure of Ferric Ammonium Sulfate (No. 21).

In the ninth put 2 measures of Cobalt Chloride (No. 27) and 1 measure of Sodium Ferrocyanide (No. 6).

In the tenth put ¼ measure of Mixed Dyes (No. 30).

Now get a large pitcher of water and fill the glasses, stirring each one slightly, after you have filled them all. The water will turn a different color as it goes into each glass. The effect is startling as well as beautiful.

EXPERIMENT 89—How to Change Water to Milk, and Milk to Ink.

Put 2 measures of Ferrous Ammonium Sulfate (No. 17) in a glass ¾ full of water. Stir to dissolve the solid. In another glass put 2 measures of Strontium Chloride (No. 10) and add a few drops of water to dissolve the solid. In the third glass put 2 measures of Tannic Acid (No. 15) and a few drops of water. Next pour the solution in the first glass into the second and the result will be milk. Pour the milk into the third glass and ink will result.

EXPERIMENT 90—Red Letters Appear When Heated.

Put 2 measures of Sodium Bicarbonate (No. 25) in a test tube. Add 2 drops of water and 2 drops of Phenolphthalein Solution (No. 22). Shake this mixture until it forms a paste-like substance. Using a clean pen, write with the mixture on a piece of white paper. The writing will probably appear slightly pink in color. When heated the writing will come out a deep red color, forming a contrast with the original faint pink.

EXPERIMENT 91—A Magic Soda Fizz.

Put 3 measures of Nickel Ammonium Sulfate (No. 13) in a test tube ⅓ full of water and warm for 2 or 3 minutes to dissolve all of the solid.

Add to the hot solution 1 measure of Sodium Carbonate (No. 4). When the action has stopped add to the solution 2 measures of Sodium Bisulfate (No. 7). The soda water is now finished, but it will not do for drinking purposes.

EXPERIMENT 92—A Column of Colors.

Put ½ inch of Sodium Silicate Solution (No. 19) in test tube and add water until the tube is ⅓ full.

In a second test tube make a solution of 2 measures of Cobalt Chloride (No. 27) and ½ inch of water. Heat the solution for a minute to completely dissolve the solid.

Pour a very few drops of the sodium silicate solution from test tube number 1 into a third clean test tube and add an equal amount of cobalt chloride solution from the test tube number 2. The two liquids together should not equal more than ½ inch in the test tube. A blue precipitate will form.

You are now finished with the cobalt chloride solution contained in test tube number 2, so that the tube may be cleaned and used for making up the next solution.

A few drops of a solution of 1 measure of Strontium Chloride (No. 10) in ½ inch of water should now be added to the blue precipitate and an equal amount of sodium silicate from the solution in test tube number 1. This will form a layer of color on top of the blue.

EXPERIMENT 93—The Disappearing—Reappearing Color.

Mix 1 measuer of Starch (No. 41) with a few drops of water, pour this mixture into a test tube full of warm water, and heat for a minute. If the solution becomes stiff on cooling add water until it is fairly fluid.

When the solution is perfectly cold add to it a few drops of starch test solution, made by adding ½ measure of Ferric Ammonium Sulfate (No. 21), ½ measure of Sodium Bisulfate (No. 7), and a few drops of Sodium Iodide Solution (No. 23) to a test tube ¼ full of water.

A deep blue color will be obtained. Now heat this blue solution to boiling and the color will entirely disappear, leaving a clear solution. On cooling the blue color will reappear. This change may be repeated as often as desired.

EXPERIMENT 94—How to Make an Unpleasant Odor.

Put 5 measures of Sodium Bisulfate (No. 7) in a test tube ¾ full of water. Shake it to dissolve the solid. Add to this solution 2 measures of Iron Sulfide (No. 43) and after a few moments a strong smell of rotten eggs will be given off.

EXPERIMENT 95—A Silicate Garden.

Put 1 teaspoonful of Sodium Silicate Solution (No. 19) and 3 teaspoonfuls of water in a clean glass. Stir slightly to thoroughly mix the liquids. Now take 1 measure of Alumnium Sulfate (No. 12), 1 measure of Nickel Ammonium Sulf e (No. 13), 1 measure of Ferrous Ammonium Sulfate (No. 17), 1 measure of Ferric Ammonium Sulfate (No. 21) and 1 measure of Cobalt Chloride (No. 27) and drop them all into the water glass solution.

Watch. In a few minutes trees, grass and all manner of chemical vegetation will begin to sprout and grow. In a short time the liquid will be filled with variously colored chemical plants and trees. The thickness of the jungle may be varied by the amount of reagents used. The glass should be set in a place where it will not be shaken.

SECTION IV

HOW TO PRESENT A CHEMCRAFT SHOW

The tricks in chemical magic described in this manual are best suited for entertainment where the audience can be seated close to the performer and his assistant.

The following directions are intended to be suggestive only and need not be followed in every instance. It is not necessary that you make-up as an Alchemist and go through all of the acting described, but appropriate patter and gestures will add very much to the interest and impressiveness of your entertainment. If possible, consider the nature of your audience before deciding on the character of your program and plan your stage setting accordingly.

Suggestions for stage settings (see Section V) are based on the program described in this Manual, Section VI the performer taking the role of an Alchemist.

PLANNING YOUR PROGRAM

Any club, group of boys or girls, or just two or three experimenters can easily stage a Chemcraft Magic Show.

The first thing to do is to decide just what tricks and experiments you are going to include in your program. You will have to decide who will be the Alchemist, who shall be the assistant, where the show will be held, and how to arrange your "theatre," stage settings, and various other details. When this is done, try all of the tricks that you are going to use, being sure to select experiments that can be done with your outfit. In the event that the experiment calls for a reagent that you do not have included in your Chemcraft Outfit it will be necessary to substitute another trick. Have your assistant work with you so that he is familiar with everything that you do. By arranging "cues" your performance will go along smoothly and without awkward pauses.

PUBLICITY

When the above details have been taken care of, you are ready to set a date and prepare the advertising for your show. If you choose, you can appointed one individual to have complete charge of publicity. In such an event he will be responsible for making the posters and tickets and to seeing that the advertising is displayed in conspicuous places.

SOUND EFFECTS

Weird sounds produced back-stage at various points during the performance will greatly add to the dramatic presentation of your show. All sound effects should be produced by your assistant.

A call bell or dinner gong which will sound a single note will prove very useful. You can have your assistant strike it back-stage upon some prearranged signal whenever you wish to arrest the attention of your audience.

SPIRIT RAPPINGS

Spirit rappings can be imitated by tapping on a half cocoanut shell with a small wooden mallet.

RUSHING SPIRITS

By having an electric fan blow against ten or fifteen inch strips of coarse paper tied together at one end and attached to the guard of the fan, a sound will be produced which suggests 'spirits' rushing through the air.

THUNDER

To produce the sound of rumbling thunder hold a piece of light sheet metal approximately 24" x 30 " along one of the 24" edges and shake it vigorously. It may be well to cover the edge to be held with adhesive tape. This will protect your fingers from the sharp edge.

Dried peas allowed to fall a short distance on to a small drum head will produce the sound of falling rain.

CRACKLING FLAMES

A sheet of cellophane (6 inches square) if crumpled and uncrumpled between the hands will produce the sound of crackling flames.

AEOLEAN HARP

By allowing the breeze from an electric fan to blow across or against a violin, guitar or any other stringed instrument, a sound resembling that of an Aeolean Harp will result.

SHRIEKING SPIRITS

The siren whistles which can be bought at any 5c and 10c store if blown unexpectedly will invariably cause shivers to run up and down the spine of your audience and produce the sound of eerie spirits in distress.

Various unusual noise-making devices are familiar to most boys and can be used to advantage when producing a magic show. However, as a usual thing, faint noises are very much more appropriate than very loud ones.

PROPERTIES

The following additional "properties" or equipment are usually needed when putting on a magic show.

1. A source of heat is necessary when doing the tricks with sympathetic inks. An ideal source is an electric toaster placed on your work table. Arrange so it can be turned on easily when needed. Be carfeul not to place any cloth or paper around the toaster as this may become ignited when the toaster is hot.
2. Two glass pitchers of one pint and one quart capacity respectively should be close at hand.
3. Ten to twenty heavy drinking glasses or plain jelly glasses which will not break easily are also necessary.
4. Unless a sink with running water is handy in the same or adjoining room, a bucket should be close at hand (behind the table) in which to pour waste liquids after completing your experiments.
5. Pleasant odors obtained by burning ordinary incense prior to the opening of the show (or during the performance) will help to heighten the air of mystery in your dimly lighted "theatre."

COSTUMES

It is highly desirable, although not necessary that the Alchemist and his Assistant be costumed for the occasion. It will make the show appear more professional and help the performers to keep from laughing while enacting their mystic roles.

ALCHEMIST'S GARB

A black or red dressing gown is suitable for the main garment. The sleeves should be wide but held at the elbow with a light rubber band. The forearm should be bare to the elbows inside the flowing sleeves of the gown. A mystic design of some sort (see illustration) should be painted on cardboard and sewed on the front of the robe. It should be made of either red, blue or black on a silver or white background.

The head should be covered with a turkish towel formed to resemble a turbin worn by a Hindu Prince or Rajah. If the Alchemist wears a long flowing robe his feet will not be visible, but if the robe does not reach to the floor the magician should wear sandals or bedroom slippers.

ALCHEMIST'S ASSISTANT—SLAVE

The assistant should be "made-up" as an Ethiopian slave. His costume can be similiar to that worn by the Alchemist (omitting the turban) or he can wear an old burlap sack with holes cut out for his head and

arms. His face and arms should be blackened with burned cork which will wash off easily when the performance is concluded. The accompanying illustration will give you a good idea of how your Slave should look when properly costumed. When addressing him, he should be referred to as your Slave. If you prefer, the blackening of the face and arms can be omitted and the Assistant can be called "Apprentice" instead of slave. By all means assign him a fantastic name such as Allah, Kola, Rota or any foreign-sounding word.

PATTER AND GESTURES

Suitable "patter" is a very important thing in every magic show. In the section following an entire program is given which contains correct "patter" to precede, to accompany and if necessary follow each trick. Become conversant with the difficult and unfamiliar words included in the patter and do not be afraid to make changes to suit your conditions and desires. The only thing to remember is, be sure that any changes you make are of an appropriate nature.

It will do no harm to have this book open before you on the table but do not make your reference to it so obvious that its presence becomes conspicious.

Always speak SLOWLY and IMPRESSIVELY with DRAMATIC pauses at the proper places. Ignore any questions, remarks or other attempts to interrupt you. Remember that your gestures must always be in keeping with the character that you have assumed. Remember that being natural will win you success and applause much sooner that ridiculous over-acting.

Upon being applauded after completing a trick, bow to the audience. Bowing can be done in one or two ways; either place the palms of each hand on opposite shoulders and bend slowly from the waist, or mesh your fingers and place your palms outward in front of your forehead when bending from the waist.

Every magician uses some type of wand with which he makes suitable gestures accompanied by appropriate patter. A wand can be athin round stick (a piece of ¼ inch dowel pin) a baton, or a piece of bessemer rod. Magicians often describe imaginary mystic symbols in the air with their wands before performing a particularly mysterious trick.

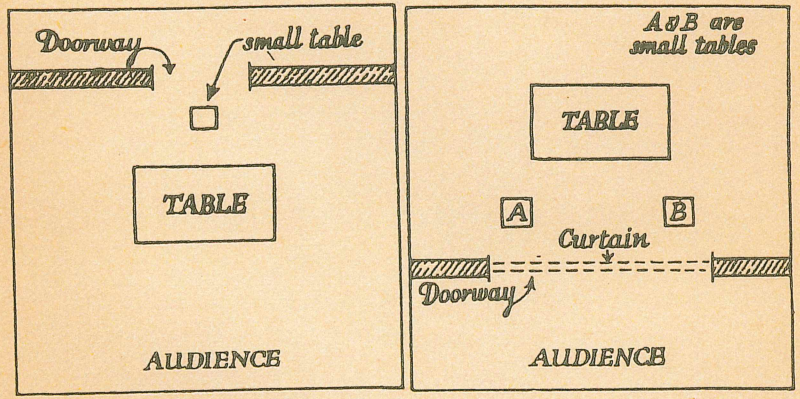
HOW TO HANDLE PERSONS WHO INTERRUPT

The best way to handle scoffers and persons who try to interrupt during the show, is to ignore them. They will soon become tired of talking to themselves and be quiet. A scoffer is usually a booster when he is talking about a trick you have just performed.

SECTION V

STAGE SETTINGS

A raised stage although advantageous is not necessary. In any case the Alchemist should have a good-sized table upon which to perform his tricks in magic. The table should be of such a height that you can work while standing behind it. Cover it with several thicknesses of newspaper to prevent any spilled liquids from penetrating to the surface. Cover the newspapers with an old piece of oilcloth, a worn out sheet, or an old tablecloth. A white cover makes it very easy for your audience to distinguish mysterious color changes which take place as you perform some of the tricks in chemical magic.



The table should be placed near a doorway and the front and sides should be draped with a cloth or piece of brown wrapping paper to conceal the bottom portion. A small table or stand placed just in rear of the large table will provide a convenient place to set tumblers, pitchers and reagents until ready for use. Figure I illustrates a way in which to arrange your stage in case you do not have an assistant, while Figure 2 shows another method which makes use of a certain effect that furnishes a "back stage" or "wings" provided you are going to have an assistant.

LIGHTING

An easy method of obtaining a dimly lighted stage (or "theatre") can be produced by using dark orange or red-tinted bulbs in the ceiling fixtures or chandelier. An additional light, such as an ordinary desk light should be handy to illuminate the beautiful color changes for your audience. You may be able to arrange a stiff paper or fiber-paper cone over the light in such a manner that white disk or light is always present on a certain portion of your white table covering. When test tubes or other vessel containing beautifully colored liquids are held in this "spotlight" they can be plainly seen against the background.

To advertise the show several attractive posters approximately 15" by 22" should be made on white cardboard. These can be painted with show card colors, water colors or lettered with crayons. Colors most suitable are blue, dark red, dark green or black. If you care to, more than one color can be used on the same poster. Information on the poster should include the name, date, place, time and price of admission together with various features of the show. To aid you in this manner a sample poster has been sketched showing a simple lay-out and the type of "ballyhoo' which might be used to advantage.

After the posters are made and placed in conspicious places (windows of houses close to the street, on a big tree in your front yard or a stake driven in your front lawn) you are ready to make the tickets.

MAKING THE TICKETS

Tickets should be cut from a sheet of light weight colored cardboard, tag board or cover stock. They should be approximately 2" x 3" and contain the necessary information to attract and direct people to your Magic Show. See the accompanying illustration. All lettering should be done with black India ink.

SALE OF TICKETS

You will do well to designate one of your friends as Business Manager. Make him responsible for the distribution and sale of tickets as well as the collection of the money. In most cases the business manager and publicity manager can be the same individual.

In deciding on an admission charge, consider approximately how many persons you will be able to seat in your "theatre." Based on this number you can determine how much you will have to charge for admission. Suppose you can seat 30 persons. If you play to a capacity house at 10c per ticket you will take in $3.00 while at 5c $1.50.

Everyone connected with the active production and management of the show should sell his share of the tickets. Have everybody go after this matter and encourage them to sell all that they possibly can.

SECTION VI

Continuity and Patter For a Complete Chemcraft Magic Show

ADVANCED PREPARATION

After arranging your theatre stage settings, lighting, a large table upon which to perform the tricks, all necessary glasses, pitchers, test tubes and reagents; and having had a dress rehearsal or two you are ready to stage your show.

PROCEDURE

After your audience has taken its seats and it is time for the "curtain to go up" have your assistant dim the lights in the theatre. Immediately following this, have him strike a dinner gong three times to prepare the audience for your entrance. Walk slowly out to the center of the stage and place yourself behind the table facing the audience. Allowing a moment to pass for applauding, and after a few moments pause, speak slowly and impressively as follows:

"Fellow mortals you should know that it is not the custom of those wise in magic to reveal their secrets to the common rabble. Nevertheless, I will today permit no ridicule or unbelief. If, therefore, be there any amongst you that is an unmightly science, Alchemy.

One word of caution I must speak. Know ye that the spirits of the Alchemists of Old, by whose aid I these wonders do perform, are proud spirits and will permit no ridicule or unbelief. If therefore, be there any amongst you that is an unbeliever in the mysteries of science, I beg that he now withdraw, lest ill results betide him here."

(Pause for an instant and then proceed as follows:)

"I now invoke the Spirits of the ancient Alchemists."

(Make several mystic gestures with your magic wand and repeat the following:)

"Quintessence of Creation, Emanation All Powerful, exert thy magic power of transformation on the unseen spirits of the dead.

Ghosts of the Past, summoned from the depths of space, I seek from out your number the Spirit of Alchemy. I bid you stay Hermes, Paracelsus, Alchemists of Old. The others may depart."

(During this recitation place one of your hands over your eyes and facing to the right appear to be struggling to gain connection with the spirit world. Remain in this position for a few moments after which your assistant should produce the effect of rushing spirits as described in Section V).

You have now arrived at the proper point to demonstrate your mystic powers and your assistant should quietly enter from off stage and select three glasses from the small stand, and carrying them on a tray, place them on the table before you. These glasses should be lightly marked 1, 2, and 3 so that you can distinguish them from each other. In No. 1 half full of water there should be placed 1 or 2 drops of Phenolphthalein Solution (No. 22. In No. 2 half full of water there should be dissolved 1 measure of Sodium Carbonate (No. 4). In No. 3 place 1 tablespoonful of clear vinegar. All of these things should be prepared beforehand so that you can immediately perform the transformation as soon as the tray containing the glasses is placed on the table before you.

Raise your wand magestically above your head and describe a magic symbol saying:

"Paracelsus the Great, they power I implore, By the incantation of this Magic Chemcraft Symbol I invoke thy aid."

Picking up the No. 1 glass in your left hand and the No. 2 glass in your right hand, and slowly pouring the contents of one glass into the other. Immediately the two colorless liquids will become the color of Burgundy Wine.

(Hold the glass in which the color change has taken place in your spotlight or before a piece of white cardboard for the audience to see the magic color. When the applause has subsided again wave your wand and picking up glass No. 3 repeat:)

"Come Spirit, and through thy magic power change this beautiful Burgundy Wine back to water."

Empty the vinegar into the glass of red liquid and immediately the glass will be filled with a colorless solution. Again hold the tumbler in the spotlight for the audience to observe the mysterious change which you have wrought.

Bow to the audience in recognition to the applause.

Immediately your assistant should enter and going to the small stand place ten glasses on a tray together with a large pitcher of water. These glasses need not be marked but should be prepared beforehand according to the following directions:

In one place 1 measure of Ferric Ammonium Sulfate (No. 21) and 1 measure of Tannic Acid (No. 15).
In the second put 3 measures of Strontium Chloride (No. 10) and 3 measures of Sodium Carbonate (No. 4).
In the third put 1 maesure of Sodium Salicylate (No. 26) and 1 measure of Ferric Ammonium Sulfate (No. 21).
In the fourth put 3 measures of Cobalt Chloride (No. 27) and 2 measures of Sodium Carbonate (No. 4).
In the fifth put 1 measure of Calcium Oxide (No. 20) and 1 or 2 drops of Phenolphthalein Solution (No. 22).
In the sixth put 3 measures of Sodium Ferrocyanide (No. 6) and 3 measures of Nickel Ammonium Sulfate (No. 13).
In the seventh put 1 measure of Ferric Ammonium Sulfate (No. 21) and 4 measures of Sodium Carbonate (No. 4).
In the eighth put ⅓ measure of Sodium Ferrocyanide (No. 6) and ⅓ measure of Ferric Ammonium Sulfate (No. 21).
In the ninth put 2 measures of Cobalt Chloride (No. 27) and 1 measure of Sodium Ferrocyanide (No. 6).
In the tenth put ¼ measure of Mixed Dyes (No. 30).

Upon having your assistant place the tray containing the glasses and pitcher of water on the table before you, address your audience as follows:

"To the true Alchemist the simple transformation which I am about to perform are mere child's play, but to you common mortals they no doubt seem mysterious and strange. See, in this pitcher I have water (lift the pitcher up). By the weird powers of the spirits of the ancient Alchemists I shall pour this water into the glasses before me, and in each produce a different color."

Pour into the various glasses as you talk, filling each about ½ full. Lift each glass in turn and hold it before the spotlight. After the applauding instruct your Slave remove the glasses and pitcher and while he is getting together the necessary materials for the next trick, address your audience as follows:

The mighty art of Alchemy for freedom stands. Behold and while the opportunity affords you may feast your eyes upon the forbidden fruit of the vine."

While the above is being recited your assistant should place 3 glasses before you on the table. These should be marked 1, 2, and 3 and should previously be prepared as follows:

In No. 1 place 3 measures of Sodium Carbonate (No. 4) and add water until the glass is ¾ full. In No. 2 place 2 or 3 drops of Phenolphthalein Solution (No. 22) and in No. 3 place 3 measures of Sodium Bisulfite (No. 5) and 3 measures of Strontium Chloride (No. 10). When the liquid in No. 1 is poured into glass No. 2 the result will be a beautiful red wine. Continue with the following patter.

"Lest some intemperate individual in your midst should too freely imbibe the fermented juices of the grape I will turn this wine to milk."

By pouring the contents of No. 2 into No. 3 glass the wine will be changed to a milky fluid.

Bow to the audience in response to their applause after which instruct your Slave to dispose of the glasses and their contents.

Instruct your Slave to bring you the necessary glasses prepared for performing the trick in which you produce patriotic colors of red, white and blue. While he is doing this recite to the audience the following patter:

"America is the land of the free and the home of the brave. 'Tis found on the freedom that soars to heights undreamed of by base mortal minds. I call upon you, O Spirits of the Alchemists, to produce the 3 colors which are the symbols of our glorious land."

This trick will require 4 glasses which are marked 1, 2, 3, and 4 respectively. Into No. 1 ⅔ full of water dissolve 1 measure of Sodium Salicylate (No. 26). In No. 2, ⅔ full of water dissolve 3 measures of Strontium Chloride (No. 10), while in No. 3, ¾ full of water dissolve ½ measure of Sodium Ferrocyanide (No. 6). In No. 4, ⅓ full of water dissolve 5 measures of Ferric Ammonium Sulfate (No. 21).

Wave your magic wand over the glasses and pour a little of the solution of No. 4 glass into each of the others thus producing the alternate colors of red, white and blue.

Acknowledge the applause by bowing twice to the audience.

The next experiment is based on the use of vinegar as a sympathetic ink. A very entertaining trick can be worked out according to the suggestions which follow.

In doing this experiment you can call upon your audience to participate in the show. The results can be made amusing as well as mysterious. It will be necessary that five or more sheets of paper be prepared previous to the show and in the manner described. All sheets should be approximately 6" x 8" and a lead pencil line should be drawn exactly through the center of the sheet. In the upper left hand corner should be placed the letter "Q." With a clean pen dipped in vinegar write such expressions or similar ones in the space provided beneath the pencil line. Be sure to use plenty of vinegar on your pen point.

1. When the seventh day of the seventh month
   Of the seventh year falls on
   Friday the 13th, BEWARE
2. Your best friends won't tell you
   But always take advice from a red bearded man
3. If at first you don't succeed,
   Look it up and be sure to read

Early to bed and early to rise
Makes you healthy, wealthy and wise

4. Brush your teeth twice a day,
Never fish during the month of May,
Write your name on a slab of clay,
Eat ice cream on your birthday

5. Never walk beneath a ladder,
Shun all evil like an adder,
Goodness gracious what does it matter?

When dry these will be invisible on the paper, but when heated the invisible writing will appear and it will not fade upon cooling.

When you come to the place in the show where you are going to use the sheets prepared as described above connect the electric toaster and addressing the audience, repeat the following:

" 'Tis not for mortal man to look beyond the bounds of time and space, but through the aid of those who dwell within the nebulous spaces of the fifth dimension I will permit you to receive messages from the great beyond.
"Oh Spirit of the ancient Alchemists, lend thy aid to those here assembled who seek secrets of the future."

Pause—and then continue as follows:

"My Slave will pass among you and distribute blank sheets of paper upon which you will write any questions concerning the future which you might wish to know. Kindly write legibly opposite the letter "Q" and I will implore the spirits of the ancient alchemists to answer them for you."

Have your assistant pass the slips among the audience and after the questions have been written have him collect them and return them to you on the stage. After he has delivered them, make several mysterious gestures over the papers as they lie before you on the table and repeat the following patter:

"Mighty Hermes, Father of Alchemy, exercise thy potent power in behalf of these poor mortals.. By the incantation of thy Servant who invokes thy aid search the future and record thy answers on these sheets of paper."

At this point pick up one of the sheets of paper and hold it against the hot toaster. Should it not prove convenient or practical to use an electric toaster the slips of paper can be heated just above the flame of your alcohol lamp. This will cause the invisible writing to appear, and after you have read it to the audience it should be returned to the writer by your Slave to keep as a souvenir of the show. Repeat the above procedure with all the slips of paper that you have at hand.

As soon as the last of the spirit messages have been read and returned to the audience you can immediately go into the next trick by repeating the following patter:

"This water into ice I will turn. Come O Spirit of Tristmegistus through thy weird powers allow thy Servant to cause this water to become ice."

Have your slave in the meantime carry from the small table and deliver to you 2 test tubes half full of liquid which have been prepared according to the following directions.

In one of the test tubes place one inch of Sodium Silicate Solution (No. 19) and add 6 or 7 drops of water. In another test tube place 8 measures of Sodium Bisulfate (No. 7) and fill ½ full of water. Shake the test tube until all of the solid is dissolved.

Upon receiving the two test tubes from your Slave hold them up for your audience to see, and then pour the contents of one into the other. Pick up your wand and make several mystic passes over and under the test tube at the same time mumbling some weird words. Continue this until the mixture gradually begins to "freeze" into transparent solid. It will become hard at the surface and gradually solidify throughout the entire mass. In a minute or two the test tube can be turned upside down and should be held in the spotlight so the audience can see that you have mysteriously turned the water into magic ice. Continue with the following patter:

"For centuries the honest tillers of the soil have attempted to learn the secrets which I am about to reveal to you, but alas, thus far their efforts have been in vain. Slave, I command that you bring me the necessary glasses that I may show these earthly mortals how easy it is to change water into milk."

At this point your assistant should transfer the 2 glasses (which have been prepared previously) from the small stand to your work table. Place in one 2 measures of Manganese Sulfate (No. 24) and in the other 2 measures of Sodium Carbonate (No. 4). Fill each glass ½ full of water.

Lift the 2 glasses from the table and hold them up for the audience to see that they contain only clear water. Set one on the table and pick up the magic wand and make several symbols over the glass. Then pick up the glass which you set down and pour its contents into the other glass. As soon as the reaction takes place repeat:

"Behold, My Friend, rich creamy milk where before was only water."

Bow to the audience to acknowledge the applause.

After your assistant has carried away the 2 glasses which were used in the last trick, instruct him to bring you a moth ball, the parcel containing 8 or 10 measures of Sodium Carbonate (No. 4) (Previously measured out), and the small olive or pickle bottle that has been filled ¾ full of vinegar. Have him place these before you on the table. Immediately drop the moth ball into the vinegar where it will drop to the bottom, then dump in the sodium carbonate from the parcel. Hold the jar up before you with your left hand, and with your right hand pick up your wand. Make several mysterious gestures over and under the bottle and repeat the following patter:

"By the power of Hermes vested in me I command that this white ball arise to the surface."

As soon as the moth ball is dropped into the liquid bubbles will form on it and after a moment carry it to the surface where it will turn around and again sink to the bottom of the vessel. At once it will begin to accummulate another covering of bubbles causing it to rise to the surface again. This action will occur several times. Time the movements of your wand to conform to the rise and fall of the moth ball.

Set the vessel containing the moth ball to one side and command your slave to bring you a handful of small, clean, dry chicken bones (such as the wish bone). These will have to be accumulated over a period of time so that you will have enough to make the trick impressive.

One of the bones which you do not show the audience until later should be kept separate from the rest, or if you prefer, mixed with the other bones but marked in such a manner that you can readily pick it out from the rest. This particular bone should be cleaned thoroughly and immersed in vinegar for 48 hours previous to your show. At the end of the first 24

hours remove it from the vinegar and test it to see whether or not it is becoming elastic. Unless it is pliable like rubber immerse it in fresh vinegar and allow it to remain there until it becomes elastic and pliable.

Select one of the bones from the pile which the Slave has placed on your table and holding it at arm's length before you, demonstrate to the audience how easily it can be broken.

To make this trick more effective you can have your assistant produce the sound of shrieking spirits (see Sound Effects in Section V) as you break each bone. Break several of them and then repeat the following patter:

> " 'Tis blasphemous these bones to break
> Excruciating pains the spirits hate
> So, the brittleness from these bones I'll take."

Nonchalantly select the prepared bone from the midst of the other bones and holding it aloft in your hand, describe the mystic Chemcraft symbol over it with your wand held in the right hand.

Then lower the bone into a tumbler ¾ full of water to which has been added 1 measure of Mixed Dyes (No. 30) and repeat such expressions as "Bula Bula, E Pluribus Unium, Hilu, Ala Baba, Ga Sunt Hite." Remove the previously made pliable bone and repeat the following:

> "I shall now demonstrate to you one of the secrets of Alchemy which has for centuries been handed down from father to son. You have just been shown that bones are brittle, and yet in the twinkling of an eye through the power of this mystic fluid, I shall before your eyes tie in this bone a knot."

Proceed to tie a knot in the bone which has been made flexible by soaking in vinegar beforehand (dipping the bone in the mixed dye solution has no effect on its pliability and is done only for effect). When the applause has died down, hand it to one of the persons in the audience and allow him to pass it among the spectators so that they can see that it is really pliable.

As the examination of the flexible bone continues, the attention of the audience should be arrested by the persistence of the spirit rappings coming from the backstage (see Sound Effects Section V). These are produced by your assistant, and when the audience becomes slightly apprehensive about the sounds, gaze toward the ceiling and with one of your hands shading your eyes and the other placed behind your ear as if listening repeat:

> "Ah! the signal from the spirit realm warning me that these revelations must cease. I feel the urge for me to return to the spirit world from whence I have derived the power to perform these mysterious feats of magic."

At this point pause and then continue:

> "However, before I depart I shall wish the air of that super agency, fire, with which my predecessors have accomplished such wonderful results, produce for you thick white clouds of magic smoke."

Having thus spoken strike a match and light your alcohol lamp, then lift your spoon into which you have placed 4 or 5 measures of Ammonium Chloride (No. 9) and with your left hand hold it over the flame of your alcohol lamp. With your right hand grasp your want and describe several mysterious gestures over the spoon until thick white clouds of smoke arise.

As soon as the smoke begins to form your assistant standing backstage

should produce the sound of rumbling thunder (see Sound Effects Section V).

As the rumbling of the thunder continues and the smoke cloud continues to belch from the spoon repeat to the audience the following final patter:

> "In thunder, in lightning, and in rain
> I do not know when we shall meet again
> I can no longer with you remain
> So must depart from this terrain."

Amidst the cheering and applause from the audience graciously bow and make a dignified exit to the dooway leading to the backstage. If after your exit the applause continues you should return and acknowledge a "curtain call" by once more bowing to the audience.

* * * * *

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For a price list of the chemicals and equipment used with this book, please refer to the Refill List in the back of any Chemcraft Experiment Book, or write to The Porter Chemical Company, Hagerstown, Maryland.

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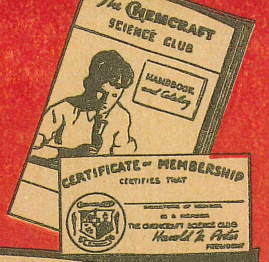

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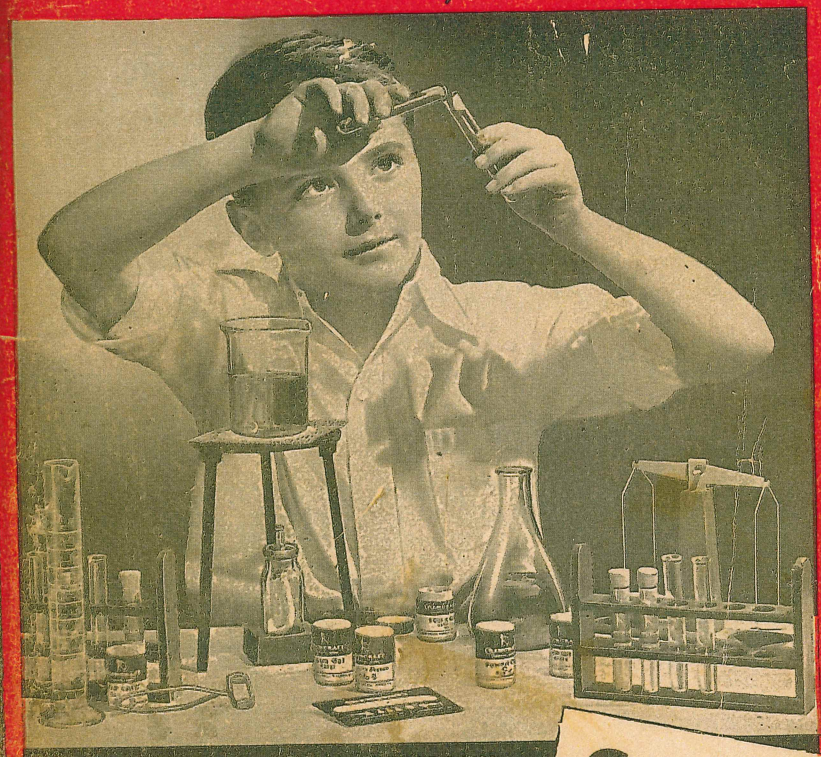

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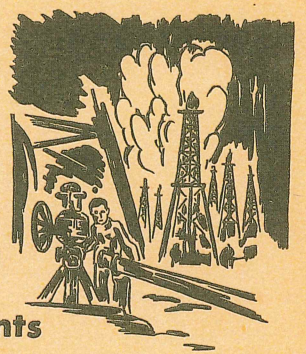

Directions for Experiments

Explanation of Chemical Science and Industries

by

Harold M. Porter

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THE PORTER CHEMICAL COMPANY
Hagerstown, Maryland

HOW TO USE YOUR CHEMCRAFT OUTFIT

CHEMCRAFT Chemistry Outfits contain no dangerous substances, but like many commonplace things they should be used in the proper way and with common sense.

By careful study and research we have prepared a great number of interesting experiments which can be performed with absolute safety if the directions in this manual are followed.

If you should eat a quantity of shaving cream it would probably make you very sick. The same is true of some of the chemicals in this outfit, although none of them are dangerous to life in the quantities supplied. If you were to strike a match while holding your finger next to the head, or ignite a whole box of matches at one time, you might be severely burned. Common sense warns you against such errors and you should exercise the same common sense when using your CHEMCRAFT. Do not combine any other chemicals, substances or materials with the chemical reagents contained in this outfit. Do not find fault with CHEMCRAFT because of incorrect results which may follow combinations of chemical compounds other than those suggested in this book. Do not expect to spill chemicals on your clothes, the furniture or the carpet without causing damage, just as happens when you spill paint, molten paraffin, milk or molasses. Always wash your hands when through experimenting, particularly after you have handled chemicals.

CHEMCRAFT should not be put into the hands of persons who are unable to read and understand the experiments in this Manual.

THE PORTER CHEMICAL COMPANY
Hagerstown, Maryland

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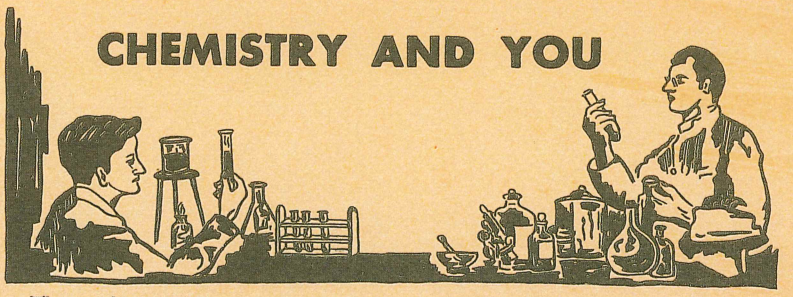
CHEMISTRY AND YOU



The entire universe about us, from its uppermost heights to its lowest depths, is built up of chemical elements and chemical compounds. Earth, sky and water are constantly passing through chemical changes. Deep in the ground, coal and oil are being formed from the remains of prehistoric forests and swamp vegetation. Precious metals and ores are being smelted under the heat and pressure of millions of tons of earth and rocks. On the surface of the earth, air and water are constantly producing chemical changes in everything they touch. All nature is merely a series of wonderful chemical reactions; plants, forests, birds, animals and people are all complex chemical engines.

Chemistry is more closely interwoven with the industries of the world than any other science, and the country which leads in chemical industries will ultimately be the richest and most powerful. It will have fewer waste materials, it will have the best manufactured articles, its food will be the most nourishing and the cheapest, and it will possess the secrets of the most powerful explosives, the hardest steels and the mightiest engines.

During the past score of years, the field of chemistry has become vastly enlarged. To illustrate how our basic industries, and we ourselves, rely on chemistry, there follows a partial list of everyday products and industries which depend directly or indirectly on chemistry for their present state of development:

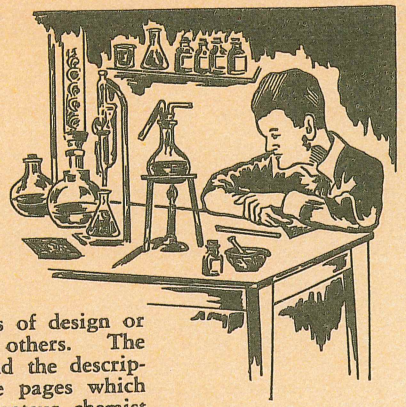
Foods	Metallurgy	Paints and Varnishes
Construction	Plastics and Rubber	Paper
Fuels	Adhesives	Petroleum
Transportation	Engineering	Refrigeration and Air Conditioning
Medicine	Glass and Porcelain	Soaps and Detergents
Electronics	Leather	Textiles
Inks and Dyes	Photography	Atomic Energy

Today, regardless of the profession a man follows, he is greatly handicapped without a knowledge of chemistry. The manufacturer, the farmer, the tradesman, the professional man, the scientist, all have constant need of chemical knowledge. In the home, the housewife is greatly handicapped if she does not have some understanding of the chemistry of the foods which she prepares and of the materials which she uses daily in her work. Mysterious chemical reactions even play an important part in enabling magicians to perform their tricks. In Chemcraft the many interesting branches of chemistry have been combined into a series of fascinating experiments which will furnish amusement for young chemists during many enjoyable and profitable hours.

The quantities of chemicals furnished with this outfit are sufficient to allow each experiment to be performed a number of times. An additional supply of any chemical or extra apparatus can be obtained by sending directly to the manufacturer. See the price list at the end of this book.

THE PORTER CHEMICAL COMPANY
Hagerstown, Maryland

HOW TO BE A CHEMIST



For a workman to be successful in any craft, there are two fundamental requirements which he must fulfill. In the first place, he must be able to use the tools of his trade with the proper skill, and in the second place, he must be able to understand and apply the principles of design or at least to follow the directions of others. The following important instructions and the descriptions of chemical apparatus on the pages which follow are designed to give the amateur chemist fundamental rules and information which will be the basis of developing skill in his craft.

Always read an experiment entirely through before starting its actual performance. By following this rule, many mistakes will be avoided.

Be sure to spread a thick layer of newspapers, oil cloth or other protecting material over the table on which you are experimenting so that liquids that may be spilled will not injure the surface.

All bottles and vials containing chemicals should be kept tightly stoppered as some chemicals become moist or gradually lose their strength when exposed to the air.

To pour a liquid from a bottle or vial, hold a glass rod against the bottom of the lip as you tip the bottle to pour from it. The liquid will then run down the rod and it is easy to control the quantity which comes from the bottle.

When using a flame, always set your candle or your alcohol lamp on a flat metal pan, such as a pie pan. This will avoid the danger of starting a fire should the candle or alcohol lamp be upset.

When you heat a liquid in a test tube, never point the mouth of it at yourself or at anyone nearby. The liquid may boil suddenly and spurt out of the tube.

When boiling a liquid in a test tube, hold the tube in a slanting position and shake or jiggle the tube slightly while heating. This will make the liquid boil much more evenly, but be careful not to slop the contents out of the tube.

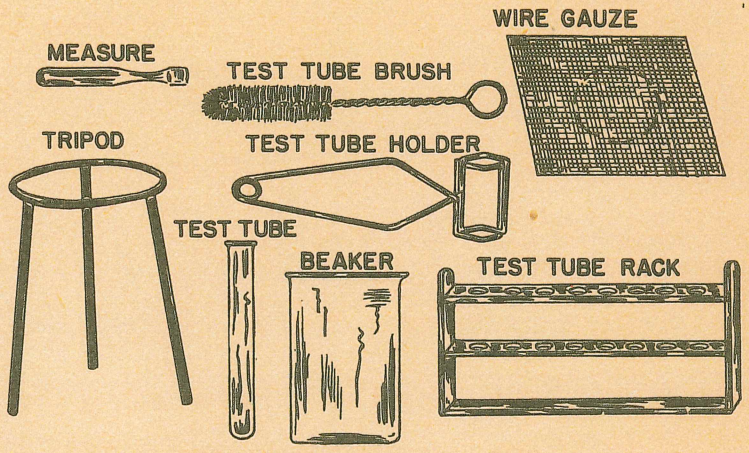
Always remove a test tube from the flame before smelling at the mouth of the tube to determine the odor of a gas.

When heating a solid in a dry test tube, warm the tube gradually. This can be done by slowly passing the test tube back and forth over the flame.

When heating a test tube over a candle, hold the bottom of the tube just above the tip of the flame. This will avoid the deposit of soot on the tube.

Cleanliness is of great importance in chemistry. Always keep your test tubes and other containers clean and wipe all of your equipment with a clean, dry cloth before putting it away.

The Use of Chemical Apparatus



THE USE OF CHEMICAL APPARATUS

This section describes the use of many pieces of chemical apparatus. Some of this apparatus may not be included in your set and the fact that it is described in this section does not mean that it is necessary for the experiments given in this manual.

MEASURE

"One measure of a chemical" means the quantity of a substance which can be held on the end of the small measure.

TEST TUBES

Test tubes are made of hard, thin glass, especially annealed to withstand heating. Liquids may be boiled and solids heated to high temperatures in test tubes, but care should be taken never to wet a dry, hot test tube, as this is almost certain to break it. When you have been heating a solid in your test tube, allow it to cool before washing it or adding a liquid.

TEST TUBE RACK

A test tube rack will be found very useful for holding your test tubes while experimenting.

TEST TUBE BRUSH

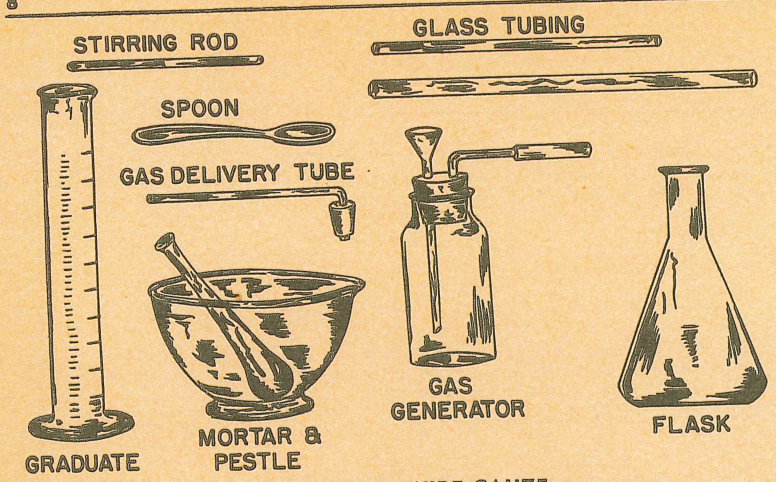
By filling a dirty tube half full of water and inserting the test tube brush, most substances can be easily removed. When test tubes are very dirty it is best to use hot water and rub the test tube brush on a cake of soap before inserting it.

TEST TUBE HOLDER

When heating substances in a test tube it sometimes becomes too hot to hold with your fingers.

BEAKER

A beaker can be used to hold hot liquids without danger of breakage. Liquids can also be heated in the beaker when supported on a wire gauze and tripod.



TRIPOD AND WIRE GAUZE

A tripod is used as a stand or support upon which you can set a vessel which you wish to heat. A wire gauze fits across the top of the tripod, making a convenient and safe stand for heating flat-bottomed vessels.

STIRRING ROD

Stirring with the glass rod often helps to dissolve a solid in a liquid. Always clean the glass rod by washing it in running water after taking it out of one liquid and before putting it into another.

GLASS TUBING

Glass tubing is used in setting up apparatus and in experiments on glass blowing.

SPOON

The spoon holds the equivalent of one teaspoonful, and is used for measuring liquids or solids. It is very convenient for heating compounds over the flame of a candle or alcohol lamp.

GAS DELIVERY TUBE

A gas delivery tube is used for the purpose of conducting a gas which is being generated in a test tube into any desired vessel.

MORTAR AND PESTLE

A mortar and pestle are used to grind or mix various substances.

GRADUATE

A graduate is used by the chemist for measuring amounts of liquids. The graduations etched in the glass wall of the vessel show you the quantity of liquid contained in the graduate.

FLASK

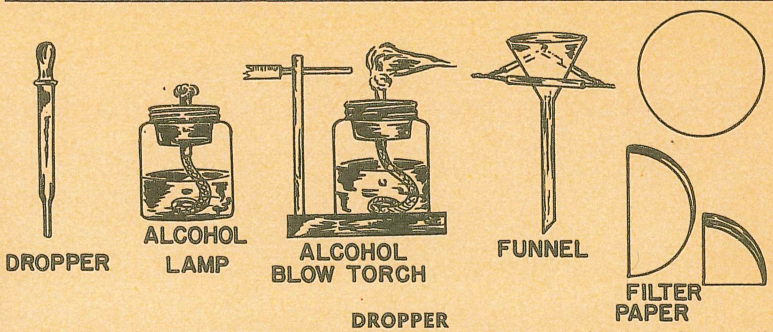
A flask will be found useful for heating liquids when the quantity is too large for a test tube. A flask is recommended when it is necessary to boil a liquid. Boiling liquids will not spurt out of a flask as easily as from a test tube. A flask should not be filled much over half full.

GAS GENERATOR

This piece of laboratory apparatus is assembled as shown in the above illustration. The funnel can be removed and used for filtering.

RUBBER TUBING

Short pieces of rubber or plastic tubing (2 inches) make excellent connectors for joining pieces of glass tubing when setting up chemical apparatus, a water bottle, or a siphon.



DROPPER

For transferring a measured number of drops of a liquid, a dropper is most convenient. Clean carefully by rinsing in clear water when you are through using it.

ALCOHOL LAMP

This is the cleanest and most convenient source of heat for most chemical experiments. Exercise great care in wiping up any alcohol that has been spilled before lighting your lamp. Do not fill the alcohol lamp near an open flame. Keep your alcohol bottle tightly stoppered and put it away as soon as you have finished filling your lamp.

ALCOHOL LAMP BLOW TORCH

With the Chemcraft alcohol lamp blow torch, it is possible to obtain quickly the high temperatures necessary in glass blowing and certain other chemical experiments. Set the alcohol lamp in the base block and insert the upright support, attaching the tube as shown in the diagram. Always be sure that the air jet tip (smallest hole) does not quite touch the flame. Adjust the air jet tip to the proper height so that when you blow into the tube, the flame is directed the greatest distance in a horizontal direction.

Blow into the tube with your cheeks puffed out, breathing through your nose. In this way you can maintain a steady stream of air from the reservoir in your cheeks. By placing the Chemcraft alcohol lamp blow torch on a suitable support, both hands are free for holding and manipulating the glass tubing or other work that is being heated in the flame.

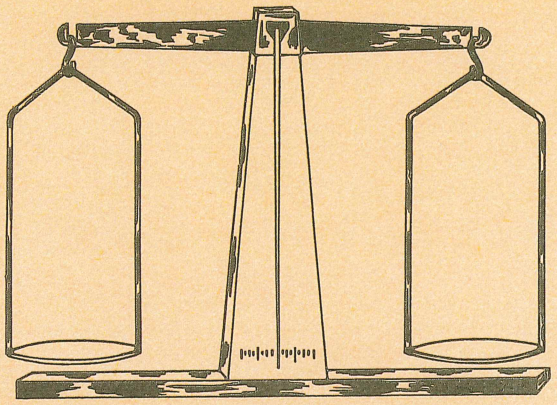
The blow torch can also be used with a candle flame where desired, by holding a candle so that the flame is in proper position with the air jet tip of the blow torch.

FUNNEL AND FILTER PAPER

It is often necessary to separate a solid from a liquid and this can usually best be done by filtering. A glass funnel and filter paper are generally used for this purpose, the filter paper being folded as shown in the diagram above. If necessary, a cone-shaped funnel can be improvised from a four-inch diameter circle of heavy wax paper. Fold this paper in half and then in half again. Cut off 1/8" at the bottom point, thus leaving a small hole at the end of the cone. Hold the wax paper in position by means of a paper clip. Now a piece of filter paper may be folded and placed into your wax paper funnel, or a small circular piece of newspaper can be used instead of filter paper. A wire triangle placed across a tumbler makes a fine support for a funnel.

BALANCE AND WEIGHTS

Balances of various types are used in every chemical laboratory. Some are used to weigh roughly large quantities of chemicals to be used in reactions; others weigh with extreme accuracy the small amounts of materials involved in chemical analysis.



BALANCE SCALE

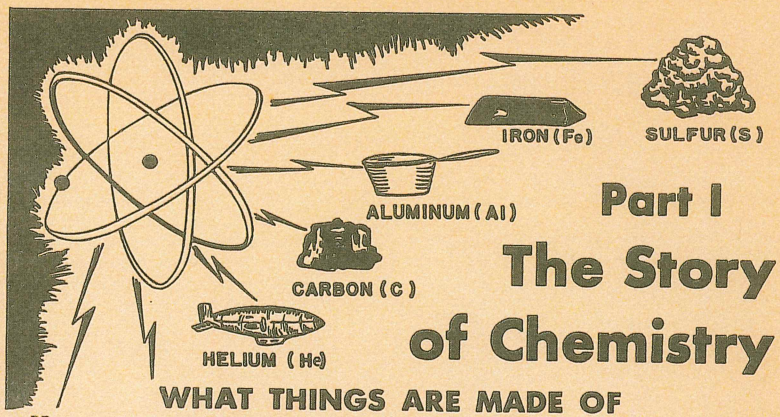
When weighing out solid chemicals, always place a piece of paper on each pan of the balance so that the chemicals will not injure the metal of the balance pan. Cut squares of smooth white paper, such as writing paper, so that the squares are of exactly the same size, and place one square on each pan of the balance. By having one on each pan, the papers counterpoise each other.

Now make sure that the beam of the balance is supported freely on its knife edges, and adjust the little metal rider on the beam so that the pointer swings an equal number of divisions on either side of the midpoint of the scale. When the pointer indicates that the instrument is properly adjusted, place the desired number of weights on the right-hand pan of the balance, and then add the chemical to be weighed to the paper on the left-hand pan until the swings of the pointer indicate that the weight of the chemical exactly equals the weights on the other pan. Remove the paper with the chemical on it from the left-hand pan.

In weighing liquids use small containers such as beakers. Place one on each pan of the balance and adjust the rider until they balance. If their weights differ considerably it may be necessary to add to one side or the other a few small weights such as coins in order to bring them into balance. When the pointer indicates that the pans are in balance, add the desired weights to the left-hand pan and pour the liquid to be weighed into the container on the right-hand pan until a balance is again reached.

In laboratory manuals and books of chemical recipes, the quantities of materials to be used are often given in terms of their weights. For example, in making the ordinary "50-50" solder, equal parts by weight of tin and lead must be melted together. You might weigh out, then, 10 grams of tin and 10 grams of lead if you wish to make 20 grams of solder.

If you need additional weights, it is easy to make them by cutting them out of sheet metal, such as aluminum or brass, and snipping them down until they just balance one or more of the weights which have been furnished. A new silver dime weighs almost exactly 2½ grams. A new nickel or 5 cent piece, weighs 5 grams. By adding several of these coins to your collection of weights, you will soon have quite a nice assortment. The gram is the unit of weight in the metric system and is used by chemists all over the world. There are about 28.35 grams in one avoirdupois ounce.



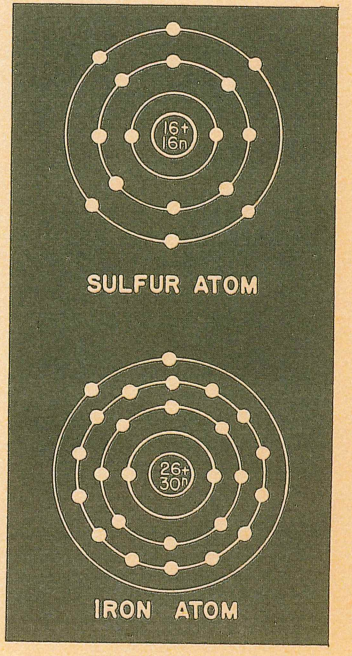
Part I
The Story of Chemistry

WHAT THINGS ARE MADE OF

Have you ever wondered of what things are made? When we look at a piece of iron or a pile of sulfur, they appear to be just that—iron or sulfur. Going more deeply, however, we find that all substances are constructed out of three kinds of units: (1) particles charged with negative electricity, called electrons, (2) particles charged with positive electricity, called protons, and (3) uncharged particles, called neutrons. All three of these units are extremely small, much too small even to be seen by the most powerful microscope.

CHEMISTRY BEGINS WITH ATOMS

These three fundamental units—protons, electrons, and neutrons—may arrange themselves in various ways to form different kinds of atoms. An atom of iron, for example, is composed of 26 electrons, 26 protons and about 30 neutrons. An atom of sulfur is composed of 16 electrons, 16 protons and 16 neutrons. Chemists have in the past been principally concerned with the atoms themselves rather than with the smaller units of which they are constructed. This is because it has been necessary to apply enormous quantities of energy (using an "atom smasher") to change the structure of an atom itself, and this could only be done on a relatively small scale. The recent development of the atomic energy piles and the production of atomic bombs have greatly increased the importance of atomic chemistry, and already the conversion of one kind of atom into another is quite commonplace to the atomic chemist.



Any substance which contains only one kind of atom is called an element. Iron and sulfur are examples of elements. An element cannot be decomposed into any simpler substance by the usual chemical processes. There are about 98 different kinds of atoms, which means the same number of elements. Some of them are very familiar to you, but many are exceedingly rare and known only to scientists.

ATOMS COMBINE TO MAKE MOLECULES

Just as protons, electrons, and neutrons join to form atoms, atoms themselves may combine with other atoms. If atoms of two or more elements combine, the result is a molecule of a compound. For example, two hydrogen atoms joining with one oxygen atom form a molecule of water, represented by the formula H_2O. To get an idea of the size of a molecule of water, suppose you could magnify a single drop of water until it was the size of our earth. The individual molecules of water in the drop would then be about the size of basketballs, but the electrons in the atoms would still be too small to see without a microscope. Several hundred thousand kinds of molecules are known, some of them built up in a very complex way out of dozens of atoms. Sometimes two or more atoms of the same element combine, forming a molecule of the element. Thus the oxygen molecule consists of two oxygen atoms.

EXPERIMENT 1—Molecules Are Very Small.

Suppose we try an experiment that will give us some idea of how tiny a molecule is and how many of them are contained in a small space.

Obtain a fairly large clear glass container, such as a quart milk bottle, a large Mason jar, or a glass pitcher. Fill this container nearly to the top with water and let it stand quietly for 5 or 10 minutes, until the water has become perfectly still in the container. Now cut with scissors a small piece of your Dye Wafer (No. 8). This should be about one-quarter inch square. Drop this dye wafer on the surface of the water in the container and watch carefully for a few minutes.

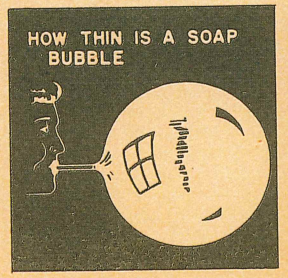
The dye wafer will float on the surface of the water and in a short time you will notice a streamer of color reaching downward through the clear water from the dye wafer. The dye is dissolving in the water and the colored molecules of the dye are diffusing through the liquid. If the water is thoroughly quiet the streamer of dye will remain concentrated until it reaches the bottom of the container, after which it will spread over the bottom and gradually diffuse upward.

After 15 or 20 minutes, shake the container slightly so that the molecules of dye are distributed evenly throughout the water. Notice that the water is tinted a decided color, showing that the small dye wafer which you used contained enough molecules to color a volume of water several hundred thousand times that of the dye.

Now empty half of the colored water from your container and refill it with clear water. Is the water still colored? Again empty half of the colored water and refill with clear water. How many times can you repeat this and still see a slight tint of color in the water? Remember that this color is caused by the colored molecules of dye and that all these molecules came from a small one-quarter inch square wafer of the dye. Does this give you an idea of how small a molecule must be?

EXPERIMENT 2—Molecules in a Soap Bubble Film.

Put a little water in a small bowl and then place a piece of soap in the water. Move the soap about with your fingers for a few minutes until some of it dissolves and the water becomes quite soapy. Now dip a bubble pipe or soda straw, or even a piece of glass tubing into the soap solution and carefully blow a soap bubble. As the soap bubble becomes larger you will notice a play of iridescent colors on the delicate film. The appearance of these iridescent colors indicates that the thickness of the bubble is only a few wave lengths of light, or only a few hundred-thousandths of an inch thick. Within this infinitesimal thickness, however, are molecules of water and molecules of soap. Just how many, we do not know, but at least this gives us some idea of how very, very small molecules must be.



MOLECULES ARE ALWAYS IN MOTION

The molecules in a substance are always moving or jiggling about. When we say that one object is hotter than another, we mean simply that its molecules are moving more energetically. When anything is heated, energy is added to it, and its molecules are made to move faster.

The molecules of a hot stove are in more rapid motion than when the stove is cool. When we touch the hot stove, its molecules jostle those of our skin, setting them, too, into abnormally rapid motion and causing us pain.

EXPERIMENT 3—The Motion of Molecules.

Fill a glass about ¾ full of water and add a teaspoonful of salt. Stir the liquid for a moment and the salt will entirely disappear. We say that the salt has dissolved in the water, which means that the crystals of salt have entirely separated and the molecules of salt being colorless, the water remains clear and of course the molecules of salt are too small to see.

Now repeat this experiment, using a glass ¾ full of fresh water, but before adding the teaspoonful of salt wait until the water becomes perfectly quiet. Pour the teaspoonful of salt into the water without stirring and the salt will settle to the bottom and will be visible there as a solid. Set the glass aside in a quiet place and let it stand overnight. The next day when you examine the glass again, much if not all of the salt will have disappeared. It has gone into solution in the water. If you hold the glass up to the light and swirl it very gently, you can see that there is a layer of dense salt solution in the lower part of the glass. Thus, although much salt has dissolved, it is not yet uniformly distributed throughout the water. There is, however, some salt even at the top, as you can easily prove by tasting a little of the water from the top of the glass. If the glass is covered and set aside for several days, the dissolved salt will slowly become distributed throughout the liquid, even without being stirred. This was largely due to the movement of the molecules of water and salt. These molecules constantly jiggling around gradually diffused the salt molecules into all portions of the glass of water.

THE ATTRACTION BETWEEN MOLECULES HOLDS THINGS TOGETHER

Molecules have an attraction for each other, that is, they tend to hang together. The soap bubble experiment which you performed is a good example of molecular attraction. Some kinds of molecules attract each other much more strongly than others. A brick, for example, is quite hard to break apart because the molecules of which it is composed attract each other strongly. So strongly do they hold on to each other that we can make them jiggle back and forth violently by heating the brick to a high temperature without forcing them to let go.

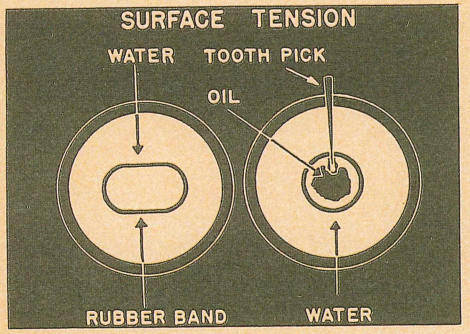
The best place to observe molecular attraction is in liquids, as all behave more or less as if their surfaces were covered with an elastic skin or film. This force, known as surface tension, is of course due to the attraction between the molecules of the liquid. It is surface tension or molecular attraction which causes a drop of liquid to assume a more or less spherical form as it falls through the air. The movement of soil moisture and of sap in plants and the concentration of ores by flotation all depend upon surface tension. Let us demonstrate this force between the molecules of a liquid with some simple experiments.

EXPERIMENT 4—The Floating Needle Shows the Attraction between Molecules.

Partly fill a saucer or small bowl with clean water. Obtain a small sewing needle, make sure that your fingers are perfectly dry, and carefully lay the needle gently upon the surface of the water. It will float there, held up by the elastic surface film due to the attraction between molecules.

EXPERIMENT 5—Demonstrating Surface Tension with a Rubber Band.

Carefully float a very light, thin rubber band on the surface of a dish of clean water that is free from any film of oil. Make sure that the band touches the water evenly all the way around. Dip a wire or toothpick into some vegetable oil, such as olive oil or salad oil, and then touch the water with the oily wire or toothpick inside of the rubber band. It may be necessary to repeat this operation several times, but you will soon notice that the band snaps out, tending to form a circle instead of its usual oval shape. This is because the oil lowers the surface tension or molecular attraction of the water inside the band, and the greater tension outside of the band then pulls on it from all directions.



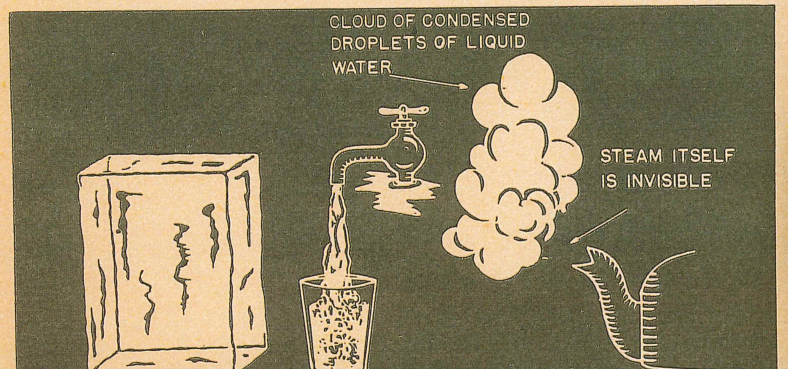
Now apply oil also to the water outside of the band. This will lower the surface tension outside so that the band springs back to its original shape.

The Difference between Solids, Liquids, and Gases

The real difference between the three states of matter - solid, liquid, and gas - is in the freedom of the molecules to move about. In solids the molecules are held in fairly definite positions, packed closely together, but their heat energy causes them to vibrate somewhat or jiggle back and forth. In liquids the attraction of the molecules for each other is not great enough to hold them firmly in place. They still stay close together, but the forces acting between them are so weak that they can slide over one another. Hence liquids can flow with ease and lack the strength and rigidity of solids. The molecules of a gas attract each other scarcely at all and hence remain far apart. Gases are, therefore, very light and have no strength or form at all.

EXPERIMENT 6—Changing a Solid to a Liquid and a Liquid to a Gas.

We have seen that when anything is heated its molecules move faster. Suppose we heat a piece of ice. Its molecules jiggle around harder and harder, until finally they jiggle right out of their places and slide around past each other. Thus the ice melts and becomes liquid. If heating the water is continued, making its molecules move still faster, every now and then one of them will be bumped so hard by its neighbor that it will sail right out through the surface and fly away among the molecules of gas which make up the air. This is why a liquid evaporates - because molecules occasionally are thrust out of it and escape. The hotter the liquid, the faster the molecules move, and the more chance they will have of escaping from the surface. That is why we warm things when we want to dry them quickly.



If water is heated strongly it will at last boil. This means that the molecules move so fast that they not only jump out of the surface very easily, but they even push each other aside in the body of the liquid and make hollow spaces - bubbles of steam, which rise through the liquid and escape. In steam, the molecules of water are whizzing about so fast that they have no chance to stick together, and hence a cloud of steam will not form a drop like liquid water, but spreads out and expands in all directions. This is why steam is so powerful for driving engines. The trillions of tiny molecules bump against the piston and push it in and out, moving trains and heavy machinery.

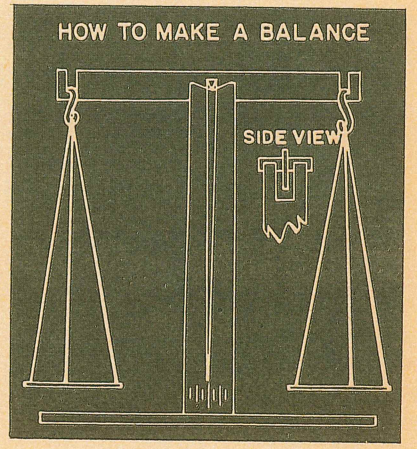
HOW CHEMISTS MEASURE MATTER

In all scientific work it is very necessary to have simple and accurate systems of measurement. In this country and in Great Britain, the English system is used for everyday measurement. This, as you know, consists of the inch, foot, and yard for linear measurement; the grain, ounce and pound for weight; and the pint, quart, and gallon for volume. In all other countries and for scientific work everywhere, the metric system is used. It is much simpler than the English system and adapts itself to scientific as well as to everyday measurements. The metric system is described and illustrated on Chemcraft Charts Nos. 2, 3, and 4.

Experiment 7—How to Make a Chemical Balance.

A chemical balance consists essentially of a base, a pedestal or upright, a beam which is pivoted at the center of the pedestal, and the pans which are suspended from each end of the beam. An indicator may be added if desired, but this is not essential.

A good size for the base is about 7″ long by 3″ wide by either ½″ or ¾″ thick. The pedestal or upright should be made out of a straight piece of wood either ¾″ square or 1″ square. This can be either 6″ or 7″ long and should be notched at the top so that the beam can be pivoted and will swing clear without touching the sides of the notch. The beam can be made of a piece of thin wood about as long as the pedestal is high, about ½″ wide, and not over ⅛ or $\frac{3}{16}$″ thick. If preferred, the beam can be made of a light, straight piece of metal in which case it will not have to be over $\frac{1}{16}$″ thick. The scale pans consist of round discs of metal or other hard material, either 1½″ or 2″ in diameter. Three strings of equal length and spaced equally around the outside diameter of the scale pans reach to the hooks which are suspended from each end of the beam.



The pivot on which the beam swings at the center can be a small, smooth iron nail. If it is desired to use a pointer, then the pivot should be triangular in shape so that the pointer can be forced onto the end of the pivot which extends to the front of the balance. In this case, small metal V's should be inserted in each side of the pedestal for the pivot to balance on. The pan hooks from the beam should be in line with the point of suspension of the beam itself, or if anything, a little below. If they are above, the beam will not swing properly, since the equilibrium at the center will be unstable.

A rider, which consists of a small piece of bent metal or even a small bent nail, is placed on the beam at the proper point to make the balance come to rest with the beam exactly level.

EXPERIMENT 8—How to Make Metric Weights.

A set of metric weights of fair accuracy can be made by using coins as standards. A new silver dime weighs almost exactly 2½ grams. Therefore, if you cut 5 equal size discs of aluminum which will exactly counterbalance a dime, each one of these discs will equal ½ gram. You can then cut larger discs which exactly counterbalance two of your ½ gram weights and these discs in turn will be 1 grams. A heavier piece of metal exactly equal to five 1 gram weights would be 5 grams, etc.

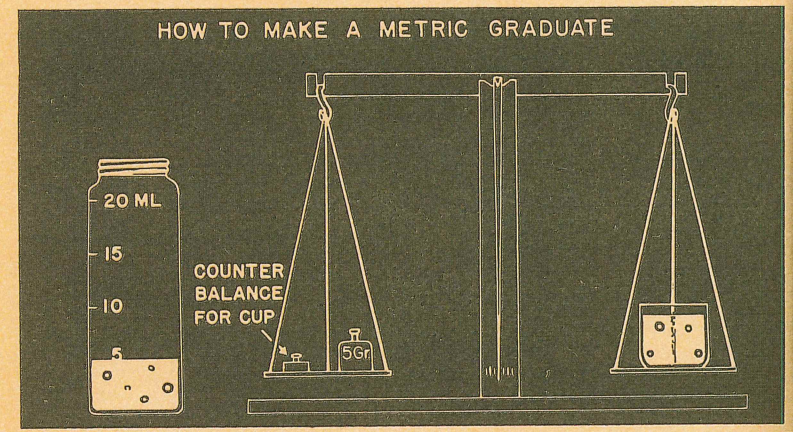
A nickel, or five-cent piece, weighs 5 grams. A copper penny weighs about 3.1 grams. A silver half dollar weighs about 12½ grams.

With your chemical balance and a set of metric weights, you will be able to weigh reasonably accurate quantities of chemicals. The desired number of weights, of course, are placed on one pan of the balance and the chemical added to the other pan until the balance comes to rest with the beam in a horizontal position.

EXPERIMENT 9—How to Make a Metric Graduate.

For this purpose you can use an empty shell vial, a tall thin olive bottle, or even a test tube.

Since 1 gram of water equals 1 cubic centimeter. or 1 milliliter, you can weigh out a certain known quantity of water and then using this volume as a standard, you can mark the graduations for your graduate with fair accuracy, using a sharp wax pencil to mark on the glass, or paste a strip of gummed paper on the glass. Use a small light cup on one side of your balance and enough weight on the other side to exactly balance the cup. The material which you use for counterbalancing does not have to be accurate weights, but can be any coins, or small bits of metal which



will exactly balance the empty cup. Now place a 5 gram weight on the side of the balance away from the cup and pour enough water into the cup to again bring the balance to rest at exactly the right position. You now have exactly 5 grams or 5 milliliters of water in your cup. By carefully transferring this water to your graduate, you can place your 5 ml. mark accurately at the bottom of the curve of the water in the graduate. Two 5 ml. portions of water will then give you 10 mls. Two 10 ml. portions of water would give you 20 mls., etc., so that your graduate can be extended to any capacity which you wish.

EXPERIMENT 10—How to Make a Metric Scale.

A meter is approximately 39-3/8". Obtain a straight, thin piece of wood and cut it exactly 39-3/8" long. Now if you divide your stick in 10 equal parts you will have decimeters. This can thus be done by dividing the length exactly in half and then, with dividers, divide each half into exactly 5 equal parts. Now divide each decimeter in turn into 10 parts. These will be centimeters and you will then have a meter scale divided into decimeters and centimeters.

HOW CHEMISTS MEASURE TEMPERATURE

On a cold winter day you say, "Brrr, it's cold. The thermometer must be near zero." Or perhaps it's summer time, mid-afternoon in August with the sun shining bright, as you wipe the perspiration from your brow

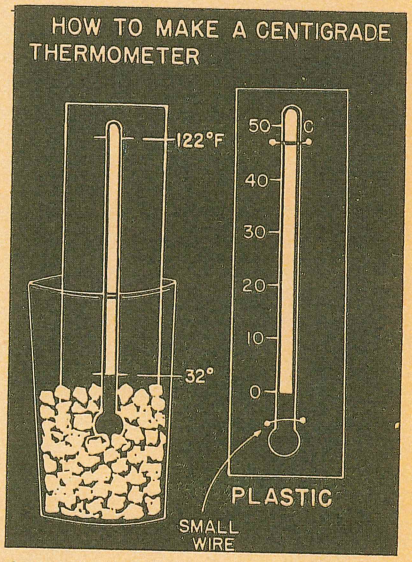
you say, "My, it must be 100° today." In both cases you are speaking of temperature in terms of the Fahrenheit scale on which the freezing point of water is 32 degrees and the boiling point of water is 212 degrees. Scientists, however, use a simpler scale for measuring temperature in which the freezing point of water is zero and the boiling point of water is 100 degrees. Chemcraft Chart No. 5 compares the Centigrade and Fahrenheit scales and shows you how to convert from one to the other.

EXPERIMENT 11—How to Make a Centigrade Thermometer.

If you have an old thermometer from which you can remove the tube, you can easily make a Centigrade thermometer, the kind which is used almost entirely in chemical work. Try to obtain a piece of plastic from which you can cut a strip about ½" wide and a little longer than the thermometer tube which you have. Drill a small hole for the hook at the top of the thermometer tube to fit in and grind out a space for the bulb near the bottom of the plastic strip. Now fasten your thermometer tube on the plastic back strip by drilling 2 sets of tiny holes and fastening the tube to the plastic by means of fine wire.

To graduate your thermometer, fill a tumbler with crushed ice, and add just enough pure water to fill the spaces between the chunks of ice. Set the thermometer so that the bulb is completely covered by the ice and let it stay in this position for at least five minutes. Now, accurately mark by means of a small scratch the exact point where the fluid within the tube stands. This will be 0 degrees Centigrade.

Now place your thermometer in a pan of water so that the bulb is again immersed and gradually heat the water to 122° Fahrenheit, adjusting the water to this temperature by means of another thermometer. Again mark the exact position of the fluid by means of a scratch on the plastic. This mark will be 50 degrees Centigrade. Now, by dividing the distance between the two marks in half and then dividing each half into 5's by means of a divider, your marks will then represent 5 degrees. If you have room, you can divide the spaces between the marks further into single degrees, but your marks in this case will be rather close together.



SPECIFIC GRAVITY

Do you float or sink in water? If you float, your specific gravity is less than one; if you sink, your specific gravity is more than one. If you neither sink nor float but just balance under the surface, your specific gravity is exactly one. Oil floats on water, therefore, it is lighter than water and its specific gravity is less than one (about 0.90). Mercury metal, of course, is much heavier than water and sinks immediately; its specific gravity is about 13.6. Specific gravity is quite important to the chemist in many ways. Chemcraft Chart No. 6 illustrates specific gravity and tells you how it is used.

EXPERIMENT 12—Specific Gravity.

Test various household articles in a pail of water to see if you can find an object whose specific gravity is equal to one. If you can find such an object, it will neither sink nor float, but will hang suspended in your pail of water between the surface and the bottom of the pail. Some vegetables have a specific gravity very

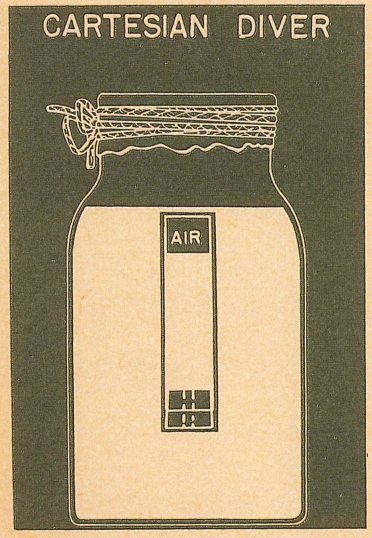
close to that of water. A fresh egg is just a little heavier than water, but a stale egg may float. Some kinds of fruit have a specific gravity very close to one. Some kinds of wood are quite light, but some of our heavier woods will actually sink in water. By driving nails into a small block of wood, it is possible to weight it so that it will have a specific gravity of exactly 1.

EXPERIMENT 13—Solution of Salt in Water Is Heavier than Pure Water.

If you have gone swimming in an ocean or in Great Salt Lake, you probably have noticed that you float better than you do in fresh water. This is because the dissolved salt increases the weight of the water. To illustrate this principle, fill two wide mouth fruit jars with water and in one of them, dissolve 3 or 4 teaspoonfuls of common table salt. Now, place an egg in the jar containing the pure water. If the egg is fresh, it will sink to the bottom of the jar. Place the same egg in the jar containing the salt solution and if you added enough salt, the egg will float on the surface. It is possible to add just enough salt so that the egg will remain suspended between the bottom of the jar and the surface of the solution.

EXPERIMENT 14—The Cartesian Driver.

Obtain a one or two quart wide mouth fruit jar. Fill the jar nearly full of water. Obtain a small vial not over 3 inches long and ½ inch in diameter. Make a hole through the cork, fill the vial about ¾ full of water and insert the cork. Now hold your finger over the hole in the cork, turn the vial upside down and lower it into your jar of water. As soon as the cork of the vial is under water, the air space in the top of the vial becomes sealed and will not change. You now want the vial to just barely float in the jar of water, that is, the bottom of the vial where the air is trapped should just barely touch the surface of the water in your jar. If the vial floats too high so that the bottom sticks above the surface of the water, remove it and add a little bit more water. If it sinks in the big jar, remove it and take out a little bit of water. Finally get it exactly balanced so that the bottom end of the vial barely touches the surface of the water in the jar.



Obtain a square of rubber, such as a piece of inner tube, a little larger than the top of the jar. Tie the rubber securely over the mouth of the jar by wrapping a piece of string around several times and pulling it tight. Now if you press slightly on the piece of rubber, you compress the air in the top of the jar and this pressure is transmitted through the water through the hole in the cork of the small vial and slightly compresses the air that is trapped in the vial. When the air in the upper part of the small vial is compressed, it occupies less space, thus decreasing the buoyancy or increasing the specific gravity of the vial causing it to sink in the jar. If your vial is balanced exactly right, you will find that you can cause it to rise or sink in the jar at will by controlling the pressure which you exert upon the rubber covering the mouth of the jar.

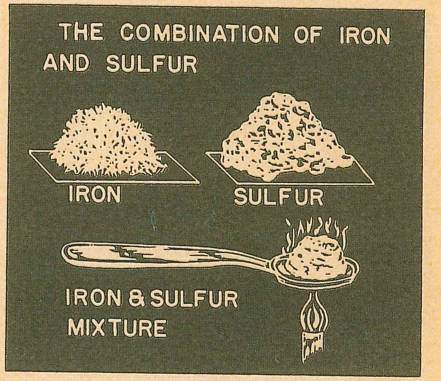
Sometimes a very small figure with a small opening at the bottom and an air space inside is weighted so that it will just float in a jar of water. Then by means of the arrangement described above, these figures can be made to rise and fall at the will of the experimenter. This figure has been called the Cartesian Diver.

THE ELEMENTS

As we have learned, the elements are the fundamental substances with which the chemist builds. Ninety-eight of them are known at present, and these are listed on Chemcraft Chart No. 1. These elements can be combined with each other in many different ways, so the number of possible chemical compounds is enormous.

EXPERIMENT 15—Two Elements Combine to Form a Compound.

Mix throughly on a sheet of paper 2 measures of Sulfur (No. 1) and 2 measures of iron filings or powdered iron metal. The powdered iron metal may consist of some steel wool cut by means of scissors into very small particles. By rolling the corners of the paper back and forth, you can mix the tiny pieces of iron thoroughly with the sulfur.



Separate the mixture into two equal parts and over one portion pass a magnet. You will find that you can pick the iron out of the sulfur in this way. The mixture can also be separated by washing the sulfur away with water, since it is lighter than the particles of iron.

Place the second portion of your iron and sulfur mixture in your spoon and heat it over a flame until it glows, or until the sulfur no longer burns. Keep your face away from the mixture while you are heating it. After the mass cools, examine it closely and see whether you can now separate the iron from the sulfur with a magnet or by washing it with water. A few stray particles of iron may have failed to combine, but most of the substance which you now have is neither iron nor sulfur, but a chemical compound of the two called iron sulfide or ferrous sulfide. You can represent the reaction which has occurred as follows:

$$\underset{\text{(iron)}}{Fe} + \underset{\text{(sulfur)}}{S} \rightarrow \underset{\text{(iron sulfide)}}{FeS}$$

Each atom of iron becomes firmly tied up with an atom of sulfur and in order to again separate them, you would have to use another chemical reaction to reverse the effect of the first one.

EXPERIMENT 16—A Compound Is Broken Down into Elements.

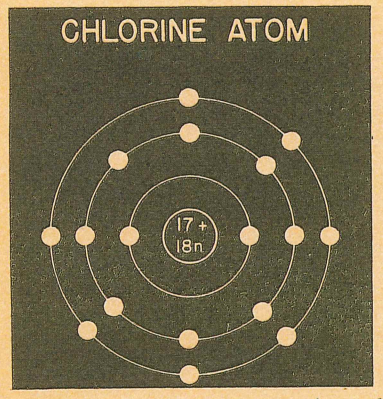
In the preceding experiment you made two elements combine to form a compound. Now let us see if we can break a compound down into its elements. Sugar is a compound. Suppose we heat 2 measures of sugar in a clean dry test tube and see what happens. Keep your face at a little distance and watch carefully the changes which take place in the sugar. Soon you will notice drops of water forming around the upper part of the test tube, therefore, sugar must contain water. Then the sugar becomes darker and finally only a charred substance remains in the test tube. This is the element carbon. Our experiment has shown that sugar is composed of water (hydrogen and oxygen) and carbon.

CHEMICAL SHORTHAND

If you will look at the list of elements, you will notice three columns after the names of the elements. The first column gives the symbol of each element. These are abbreviations and are derived from the first letter or the first and a subsequent letter of either the English or Latin name of the element. The chemical name for common table salt is sodium chloride. The chemist would write NaCl, for natrium, the Latin for sodium, and Cl for chlorine.

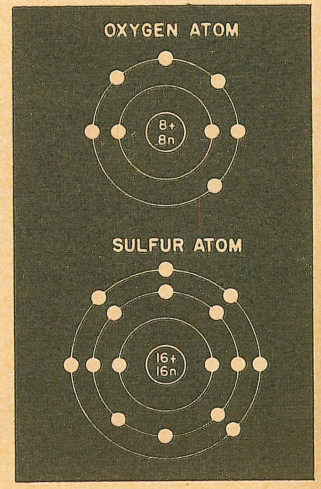
HOW ATOMS ARE CONSTRUCTED

In order to understand how elements combine to form compounds and how compounds react with each other, we must return to the atom and see how it is constructed. An atom consists of a nucleus, or heavy center portion, around which are gathered very light electrons. The nucleus is made up of positively charged protons and uncharged neutrons. If we could enlarge an atom ten billion times, it would be about a yard across. Even with this tremendous magnification, the nucleus would be only a tiny pinpoint, and the individual protons and neutrons could probably not be seen. Small as it is, this tiny nucleus contains almost all the weight of the atom. Each of the electrons is about the size of the nucleus but weighs practically nothing. There are as many electrons outside the nucleus as there are protons in the nucleus, and they are believed to revolve about the nucleus very much as the planets of our solar system revolve around the sun. An atom is mostly empty space, somewhat like a swarm of mosquitoes, since the electrons are relatively far apart.



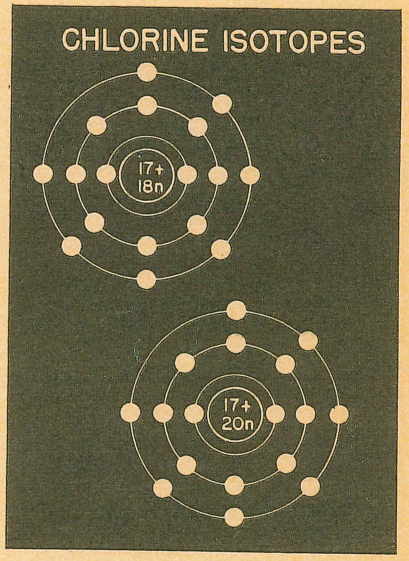
SOLVING THE MYSTERY OF ATOMIC WEIGHTS

In Chemcraft Chart No. 1, the List of Elements, you will find a column headed "Atomic Weights". Of course, an atom is entirely too small to weigh directly, but nevertheless chemists have been able to work out the relative weights of atoms of all the elements. To serve as a basis of comparison, the atom of oxygen was assumed to have a weight of 16 units, and the weights of the other atoms were compared with it by analyzing large numbers of compounds to find out the ratios by weight in which each of the elements combines with oxygen. Then, making allowance for the numbers of the different atoms in the compounds, it was possible to find the relation between the weight of each kind of atom and that of the oxygen atom. For example, we find by experiment that the compound sulfur dioxide contains equal weights of sulfur and oxygen. Now we know that in this compound there are only half as many atoms of sulfur as oxygen, and hence each sulfur atom must be twice as heavy as an oxygen atom if all together they are to weigh as much as double their number of oxygen atoms. Thus, letting the weight of the oxygen atom be 16, the relative atomic weight of sulfur is 32. Atomic weights are important because from them we can compute the relative weights of elements which enter into reactions.



WHAT ATOMIC NUMBERS MEAN

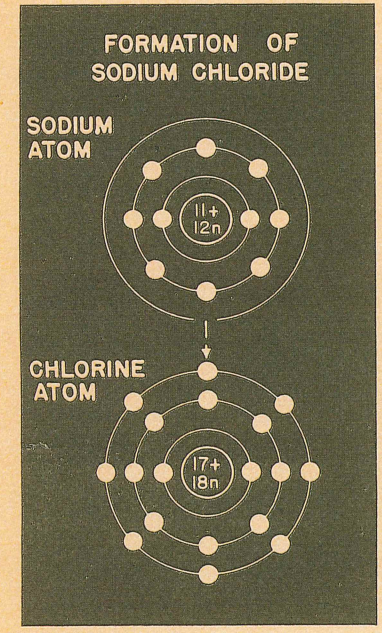
The column headed "Atomic Numbers" in Chemcraft Chart No. 1 tells us the number of protons in the nucleus of each kind of atom. These numbers are deduced by studying the x-rays emitted when each element in turn is made to serve as the target of an x-ray tube. Since the protons each carry one unit positive charge, their number tells us the positive charge of the atomic nucleus, and the number of electrons surrounding the nucleus must be exactly equal to this in a normal atom. Consider an atom of chromium. Its atomic number tells us there are 24 protons in its nucleus and hence 24 electrons around its nucleus. But its atomic weight is 52. Each proton has a weight of about one unit, and the electrons weigh almost nothing. Therefore, we know that the nucleus of chromium must contain enough neutrons (each with a weight of one unit) to build the weight up to 52. We picture, then, the chromium nucleus as composed of 24 protons and 28 neutrons. Try the same reasoning on the oxygen atom and see whether you come out with 8 protons and 8 neutrons. Many of the atomic weights are not even close to whole numbers. This simply means that not all the atoms of these elements have the same weight. Ordinary chlorine, for example, is a mixture of atoms having weights of 35 and 37. The atomic number of both kinds is 17, one contains in its nucleus 17 protons and 18 neutrons, and the other contains 17 protons and 20 neutrons. These two kinds of chlorine atoms are called isotopes. Some elements have as many as a dozen isotopes.



HOW ATOMS FORM IONS

The attraction of the positive nucleus for the negative electrons surrounding it binds the atom together. Certain atoms, however, may tend to lose one or more of their outermost electrons. Thus the hydrogen, sodium, and potassium atoms tend to lose one electron each; the calcium, magnesium, and copper atoms tend to part with two electrons; and the aluminum atom may give up three electrons. On the other hand, certain other atoms are so constructed that their tendency is toward gaining additional electrons rather than losing some which they already possess. Thus, the chlorine, bromine, and iodine atoms tend to gain one electron apiece, while the oxygen and sulfur atoms may gain two. Of course an atom which has gained or lost electrons is no longer a normal, neutral atom. An atom which has more than its normal number of negative electrons is negatively charged and is called a negative ion. One which has lost electrons is left with a positive charge and is called a positive ion.

HOW ATOMS CAN COMBINE



Suppose we bring together an electron-losing atom, such as sodium, and an electron-gaining atom, such as chlorine. An electron will then leave the sodium atom and go over to the chlorine atom. The sodium atom, Na, becomes a sodium ion, Na^+, and the chlorine atom, Cl, becomes a chloride ion, Cl^-. These ions, having opposite electrical charges, attract each other and tend to cling together, forming units in a solid crystal of sodium chloride, common salt, NaCl. It would, of course, be more accurate to write its formula Na^+Cl^- but we rarely bother to do so. A positively charged ion is called a cation and a negatively charged ion is called an anion.

When a substance is dissolved in water, it can react much more readily with other substances because the molecules of which the compounds or elements are composed are free to move about in the solution and combine with each other.

EXPERIMENT 17— To Show That Water Aids Chemical Reaction.

Mix thoroughly on a sheet of paper 1 measure of Strontium Chloride (No. 10) and 1 measure of Aluminum Sulfate (No. 12). Notice that there is no change in these chemicals and you can see no sign of any chemical reaction.

Now dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube ¼ full of water and dissolve 1 measure of Strontium Chloride (No. 10) in another test tube ¼ full of water. Pour these solutions together and you will immediately see a heavy white precipitate showing that the chemicals have reacted and have formed a new compound. This is because the ions which made up the crystals of strontium chloride and aluminum sulfate were separated from each other when the compounds were dissolved in water. They were thus free to enter into a new combination, forming the insoluble white compound, strontium sulfate.

EXPERIMENT 18—Another Chemical Reaction Depending on Ionization.

Mix 1 measure of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21) on a piece of paper and see if you can see any signs of a chemical reaction. Now pour the mixture of dry chemicals into a test tube half full of water and note that a chemical reaction immediately takes place, due to the formation of new compounds. Here again the ions have separated when the compounds are dissolved, and are then able to enter into a new combination.

WHAT IS AN ELECTROLYTE?

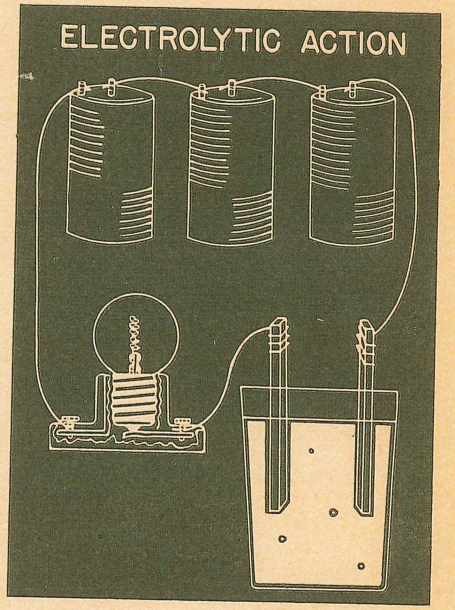
When a solid salt, for example sodium chloride, which is composed of sodium ions and chloride ions, is dissolved in water, the water molecules pry apart the ions, overcoming their attraction for each other. Thus the ions are free to float away in the liquid, and the solid salt crystal dissolves. A salt solution, therefore, contains no molecules of sodium chloride but only sodium and chloride ions, although for the sake of simplicity we referred to them in an earlier experiment as molecules. Such a solution, since it contains electrically charged ions, is a conductor of electricity and if wires leading from the poles of a battery are dipped into it, a current flows, carried through the solution by the ions.

Many other compounds resemble sodium chloride in general structure. The typical salts, such as sodium bromide, potassium nitrate, sodium phosphate, copper sulfate, barium chloride, etc., are composed of ions, as are bases, such as sodium hydroxide and calcium hydroxide. Acids in the pure state usually are not made up of ions, but as soon as they are dissolved in water a reaction occurs which produces many or few ions, depending upon whether the acid is strong or weak. Compounds like those we have been discussing are called electrolytes, since their solutions contain ions and hence conduct electricity.

EXPERIMENT 19—An Electrolyte Will Conduct Electricity.

Connect 3 dry cell batteries in series and hook a 3 volt flashlight bulb to one of the wires coming from the dry cells. Fasten a small zinc or copper plate, or a carbon rod, to the free end of each wire, as shown in the accompanying drawing. Now, fill a tumbler or a jelly glass about half full of water and put your plates or carbon rods in the glass so that they are partly immersed in the water. Move them back and forth, varying the distance between them and notice that your flashlight bulb will not light unless the plates or rods actually touch each other.

Now dissolve 1 or 2 teaspoonfuls of common table salt, sodium chloride, in the water and again introduce the plates or rods which are connected with your batteries. In this case, you will find the flashlight bulb will light even though the plates do not touch each other, showing that electric current is going through the solution.



ANOTHER WAY IN WHICH ATOMS CAN COMBINE

Sometimes atoms combine with each other in a different way. Instead of one atom giving up completely an electron to another, two atoms may share a pair of electrons between them, one from each atom. Since both atoms are trying to hang on to the same pair of electrons, they stick together, forming a molecule of the compound. Some examples of this type of combination are water, H_2O; carbon dioxide, CO_2; alcohol, C_2H_6O; glycerine, $C_3H_8O_3$; and sugar, $C_{12}H_{22}O_{11}$. Such compounds are not composed of ions, and provided no reaction occurs when they dissolve, their solutions contain no ions and hence do not conduct electricity.

EXPERIMENT 20—Showing that Sugar is not Ionized.

Dissolve a teaspoonful of sugar in a glass half full of water, and, using the apparatus for the experiment on electrolytes, prove that a sugar solution will not conduct electric current.

CHEMICAL EQUATIONS

Now that we have learned how atoms form ions and combine with other ions to form molecules, this is a good place to find out how the chemist expresses these reactions by means of equations.

Before you write an equation for a reaction, you must know what substances react and what substances are formed by the reaction. Suppose you add sulfuric acid, H_2SO_4, to sodium bicarbonate (baking soda), $NaHCO_3$, and obtain as products carbon dioxide gas, CO_2, sodium sulfate, Na_2SO_4, and water. First write down the formulas for the reacting substances and the products:

$$NaHCO_3 + H_2SO_4 \rightarrow CO_2 + Na_2SO_4 + H_2O$$

Notice that in order to form a molecule of sodium sulfate you must put in two sodium atoms. To get two sodium atoms you need two molecules of sodium bicarbonate. Therefore, on the lefthand side of the equation write: $2\ NaHCO_3 + H_2SO_4$. Having taken two molecules of sodium bicarbonate, you know that you are going to get out two molecules of carbon dioxide on the righthand side, and if you count up the hydrogen and oxygen atoms, you will find that enough are left over to form two molecules of water. The balanced equation, then, is:

$$2\ NaHCO_3 + H_2SO_4 \rightarrow 2\ CO_2 + Na_2SO_4 + 2\ H_2O$$

No atom is ever lost in a chemical reaction. Atoms may be rearranged, re-shuffled, and re-dealt, but they are never destroyed and never created in ordinary chemical processes. This is the law of conservation of matter.

Chemical equations are simply concise ways of writing down what one actually finds out about a reaction by trying it. It must not be supposed that every nicely balanced equation that can be written will always mean something. It is easy to think up all sorts of reactions which look well on paper but do not actually work in the laboratory. The advanced branches of chemistry called thermodynamics and chemical kinetics tell us which reactions will actually react and which will not, but usually the easiest way to find out is to try the experiment and see.

METALS AND NON-METALS

The elements can be more or less definitely divided into two classes — metals and non-metals. The metals are the elements whose atoms lose negative electrons, forming positive ions. Atoms of non-metallic elements tend rather to gain electrons, forming negative ions, or at least to acquire a share in some of the electrons of another atom. Some of the elements are hard to classify in this way since their atoms may either gain or lose electrons depending upon circumstances. Chemically, such elements have both metallic and non-metallic properties.

ACIDS AND BASES

In general, oxides (formed when metallic elements combine with oxygen) unite with water to form bases. Bases are compounds containing hydroxide ions, OH^-. When they are sufficiently soluble, their solutions turn red litmus paper blue. Most of the oxides of non-metallic elements, on the other hand, unite with water to form acids. The solutions of acids contain hydrogen ions, H^+, and turn blue litmus paper red.

EXPERIMENT 21—Making Sulfurous Acid.

Put 2 measures of Sulfur (No. 1) in your spoon and heat the spoon over a flame until the sulfur takes fire and burns. Now rest the spoon containing the burning sulfur on a metal plate and hold over it, mouth down, a one-pint milk bottle, so that the fumes from the burning sulfur will go into the bottle. After you have trapped some of the fumes, cover the mouth of the bottle with your hand and carefully pour about ½ inch of water into the bottle. Again close the mouth of the bottle tightly with your hand and shake for a moment so that the fumes within the bottle can be dissolved by the water. Now moisten a piece of blue Litmus Paper (No. 18) with the water from the milk bottle. It will turn red, showing that you have formed an acid. Sulfur, when burned in air, produces sulfur dioxide, which is an oxide of a non-metal. When this is dissolved in water, we obtain sulfurous acid.

$$\underset{\text{(sulfur dioxide)}}{SO_2} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(sulfurous acid)}}{H_2SO_3}$$

EXPERIMENT 22—Making Carbonic Acid.

Put 2 measures of baking soda in a test tube, add a few drops of vinegar and immediately place a gas delivery tube and stopper in the test tube. Fill another test tube half full of clear water and allow the gas being generated in the first tube to bubble through the water in the second tube. After the gas has bubbled through the clear water for 2 or 3 minutes, test the water in the second tube with a piece of blue Litmus Paper (No. 18). You will find that the blue litmus paper turns slightly pink, showing that you have a weak acid.

When you added vinegar to baking soda, carbon dioxide gas was given off and when carbon dioxide gas is dissolved in water it forms a weak acid known as carbonic acid.

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(carbonic acid)}}{H_2CO_3}$$

EXPERIMENT 23—Making a Base.

Put 2 measures of Calcium Oxide (No. 20) in a test tube and fill the tube about half full of water. Shake thoroughly to dissolve as much of the solid as possible. After the remaining solid has settled, test the clear liquid in the tube with a piece of red Litmus Paper (No. 18). The litmus paper will turn blue, showing that you have a base. You dissolved calcium oxide, which is an oxide of a metal, in water, a process which we have learned forms a base.

$$\underset{\text{(calcium oxide)}}{CaO} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(calcium hydroxide)}}{Ca(OH)_2}$$

WHAT WE MEAN BY THE pH OF A SOLUTION

It is frequently necessary to state just how acidic or basic a given solution is. Of course, this depends upon the quantity of hydrogen ions or of hydroxide ions in a unit volume (such as a liter) of the solution, but for many purposes it is inconvenient to state this quantity itself. Instead we derive from it by mathematical means a number called the pH of the solution, which tells at a glance how strongly acidic or basic the solution is. Chemcraft Chart No. 7 shows that the pH of pure water is 7, while the acidic solutions have lower pH values and the basic or alkaline ones have higher pH values.

The chart summarizes also the important uses of the pH system in chemistry, biology, medicine, engineering, and agriculture.

SALTS

Salts are compounds made up of positive and negative ions but containing neither hydrogen nor hydroxide ions. Common table salt, sodium chloride, is the most familiar example, but there are hundreds of others. A salt is formed whenever an acid and a base neutralize each other. For example, sodium chloride is formed when hydrochloric acid and sodium hydroxide react. Similarly, calcium sulfate is produced by the reaction of calcium hydroxide with sulfuric acid.

EXPERIMENT 24—Making a Salt.

Put 2 measures of baking soda in a test tube, add a few drops of vinegar and immediately place a gas delivery tube and stopper in the test tube. Fill another test tube half full of clear water and allow the gas being generated in the first tube to bubble through the water in the second tube. After the gas has bubbled through the clear water for 2 or 3 minutes, set this tube to one side.

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(carbonic acid)}}{H_2CO_3}$$

In another test tube place 1 measure of Calcium Oxide (No. 20) and fill the tube half full of water. Shake thoroughly to dissolve as much of the calcium oxide as possible and after the excess has settled to the bottom of the tube, pour the clear solution into another test tube.

$$\underset{\text{(calcium oxide)}}{CaO} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(calcium hydroxide)}}{Ca(OH)_2}$$

You now have in the first test tube carbon dioxide gas dissolved in water. This is a non-metallic oxide dissolved in water, which we have learned forms an acid; in this case carbonic acid. In the second test tube you dissolved a metallic oxide in water which formed a base, calcium hydroxide. Now pour the contents of one of the tubes into the other and note that you obtain a white precipitate. You now have added an acid to a base, forming a salt.

$$\underset{\text{(carbonic acid)}}{H_2CO_3} + \underset{\text{(calcium hydroxide)}}{Ca(OH)_2} \rightarrow \underset{\text{(calcium carbonate)}}{CaCO_3} + \underset{\text{(water)}}{H_2O}$$

When a strong base, such as sodium hydroxide, combines with a weak acid, such as carbonic acid, the salt which is formed is basic or alkaline in reaction. On the other hand, when a strong acid, such as sulfuric acid, combines with a weak base, such as aluminum hydroxide, the resulting salt gives an acid reaction.

Another type of salt which may give an acid reaction in solution is the so-called "acid salt", formed when an acid is treated with too little base to neutralize it completely. An example of an acid salt is sodium bisulfate, formed when we add to sulfuric acid only half the amount of sodium hydroxide which would be required to entirely neutralize the acid.

EXPERIMENT 25—An Alkaline Salt.

Fill a test tube half full of water and add 1 measure of Sodium Carbonate (No. 21). Shake the tube for a moment to completely dissolve the solid. Now cut a small strip of red Litmus Paper (No. 18) and dip this into the test tube containing the sodium carbonate solution. The litmus paper will turn blue, showing that the sodium carbonate gives a basic or alkaline solution.

EXPERIMENT 26—An Acid Salt.

Again fill a test tube half full of water and add 1 measure of Sodium Bisulfate (No. 7). Cut a small piece of blue Litmus Paper (No. 18) and moisten this with the sodium bisulfate solution. The blue litmus paper will turn red, showing that sodium bisulfate gives an acid solution.

EXPERIMENT 27—A Neutral Salt.

Place 1 measure of sodium chloride (common table salt) in a test tube half full of water. Shake until the solid is entirely dissolved. Now test this solution with first red and then with blue Litmus Paper (No. 18). Does the color of either piece of litmus paper change? If not, this shows that sodium chloride is a neutral salt.

EXPERIMENT 28—Testing Vinegar.

Is vinegar an acid or a base? Test some vinegar with red and blue litmus paper, keeping in mind that if the blue litmus paper turns red, it is acid and that if red litmus paper turns blue, it is a base.

EXPERIMENT 29—Testing Lemon Juice.

Test a few drops of lemon juice in the same manner that you tested the vinegar. Does lemon juice contain an acid or a base?

EXPERIMENT 30—Testing Ammonia Water.

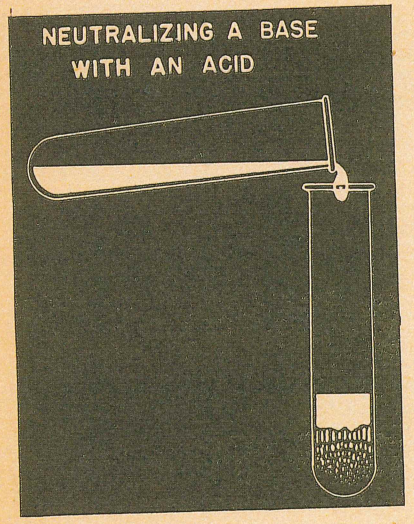
Test a few drops of household ammonia with red litmus paper and blue litmus paper. Is ammonia an acid or a base?

EXPERIMENT 31—The Reaction of an Alkaline Salt with an Acid Salt.

Fill a test tube ¼ full of water and add 1 measure of Sodium Carbonate (No. 4). Shake to dissolve the solid. In another test tube ¼ full of water, dissolve 1 measure of Sodium Bisulfate (No. 7). Now pour the contents of one test tube into the other. The liquid fizzes up just like soda water. As we have learned, sodium carbonate gives a basic solution and sodium bisulfate an acid solution. When we brought them together, therefore, they tended to neutralize each other, forming sodium sulfate and liberating carbon dioxide.

EXPERIMENT 32—Neutralizing a Base with an Acid.

Place ½ measure of Calcium Oxide (No. 20) in a test tube half full of water and shake the tube well to dissolve as much of the calcium oxide as possible. As you have already learned, calcium oxide reacts with water to form calcium hydroxide, and so you now have a solution of calcium hydroxide. Set the test tube aside so that the undissolved solid can settle out.



In another test tube place 2 measures of Sodium Bisulfate (No. 7), fill it half full of water, and shake the tube until the solid is dissolved.

Now take a third clean test tube and pour into it some of the calcium hydroxide solution. The solid particles should have settled by this time, and if you are careful you can pour off just the clear part. Add to the clear solution 1 drop of Phenolphthalein Solution (No. 22) and notice that a clear red color appears, showing that the solution is alkaline, since the color change is a test for the presence of an alkali or base.

Next add some of the sodium bisulfate solution, a few drops at a time, to the calcium hydroxide solution; and if you work carefully, you will find that the red color becomes fainter until finally one more drop of sodium bisulfate solution, which is acid, will turn the phenolphthalein colorless. You now have neutralized the alkaline calcium hydroxide solution by adding acid, and the solution which you have in the test tube is now neither alkaline nor acid (at most, slightly acid). It is a neutral solution. The salt formed in this case is calcium sulfate.

INDICATORS

There are certain substances known as indicators, which have the property of turning one color in the presence of acids and another color in the presence of bases. They can thus be used to show whether any liquid is acidic, basic, or neutral.

There are many different kinds of indicators, and they are very interesting substances. The following experiments will introduce you to some of the common indicators.

EXPERIMENT 33—Litmus Paper.

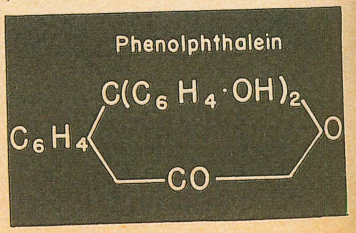
Put a piece of blue Litmus Paper (No. 18) into a glass half full of water. Add ½ measure of Sodium Bisulfate (No. 7) (or a few drops of vinegar) and stir. Next add 2 measures of Sodium Carbonate (No. 4) and stir again. Notice the color change.

EXPERIMENT 34—Phenolphthalein.

Put 1 or 2 drops of Phenolphthalein Solution (No. 22) into half a glass of water. Drop in one measure of Sodium Carbonate (No. 4), which is an alkali, and stir. Next drop in 2 or possibly 3 measures of Sodium Bisulfate (No. 7), which is an acid, or a few drops of vinegar (which contains acetic acid) and stir again.

Phenolphthalein is an indicator which is red in the presence of alkalies and colorless in the presence of acids. Some water supplies are alkaline and hence give the red color as soon as the phenolphthalein is added.

The molecule of phenolphthalein is a complicated structure built up of atoms of carbon, hydrogen, and oxygen joined together as shown in the diagram.



EXPERIMENT 35—Phenolphthalein Paper.

Dip a piece of soft, white paper into Phenolphthalein Solution (No. 22) and spread it out to dry. When it is dry, you have a piece of paper which you can use as an indicator merely by moistening it. Place ¼ measure of Sodium Carbonate (No. 4) on the paper and add a drop of water. Notice the red color which forms immediately.

Moisten a piece of the phenolphthalein test paper and hold it over the mouth of a bottle of household ammonia. It will turn red, showing that ammonia fumes are alkaline.

EXPERIMENT 36—Testing a Laxative Tablet for Phenolphthalein.

Obtain a stick of laxative chewing gum or a laxative tablet and break it up into small pieces. Put these into half a glass of water and stir for a minute or two to extract some of the phenolphthalein. Now add one measure of Sodium Carbonate (No. 4) and if the tablet contains phenolphthalein, the liquid will turn red.

Many of the most common flowers, fruits, and vegetables with which you are familiar contain coloring matter which belong to a class of indicators. Try some of the following experiments if you can obtain the necessary flowers, fruits, or vegetables.

EXPERIMENT 37—Hollyhock Indicator.

The purple hollyhock contains a coloring matter which makes a very good indicator. Break up several purple hollyhock flowers and place them in a tumbler or small dish. Now add to them about ½ glass of hot water and allow them to stand with occasional stirring until the liquid becomes a deep purple.

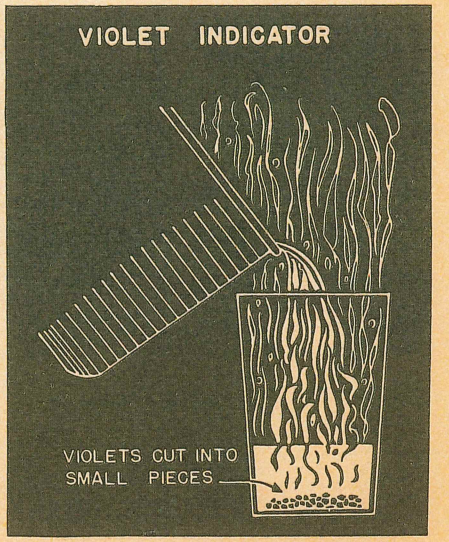
EXPERIMENT 38—Hollyhock Test Paper.

Moisten a piece of soft, white paper with the purple hollyhock liquid and when it is dry, use it to test for acids or alkalies. This paper is called "Mallow Paper" and it is frequently used as a test paper because it is quite sensitive to the presence of either acids or alkalies.

When you make these different test papers, it will be a good plan to label them so that you know what each one is later on.

EXPERIMENT 39—Violet Indicator.

Violets contain a coloring matter that they may be used as an indicator. Suppose you pick about two dozen of these flowers, remove the stems and cut the blossoms into tiny pieces. Now extract the coloring matter with hot water as you did in Experiment 37. When you obtain a good greenish-blue solution, test a small portion with some Sodium Carbonate (No. 4) dissolved in water, and then with some vinegar. Notice the sharp color changes. The coloring matter in violets is red in acids, and green in alkalies.



EXPERIMENT 40—Violet Test Paper.

To make violet indicator test paper, moisten a piece of soft, white paper with your violet liquid prepared in the preceding experiment. When your paper is dry, test it with an alkali (sodium carbonate) and an acid (vinegar). Here again you will notice that the color changes as it did in the experiment on a violet indicator.

EXPERIMENT 41—Flag Indicator.

The blue flag or iris contains a coloring matter which may be used as an indicator. If you can obtain some blue flags, put them in a glass and extract the coloring matter with hot water. When a deep blue solution is obtained, test a small portion of it with a little sodium carbonate and then with a little sodium bisulfate or vinegar and notice the color change.

EXPERIMENT 42—Dahlia Indicator.

If you can obtain some purple dahlias, a very good indicator can be made from them. Break the flowers into small pieces and pour half a glass of hot water over them. When the solution has become a deep purple, test a small portion of it with ¼ measure of Sodium Carbonate (No. 4) and also with a little acid.

The coloring matter contained in purple dahlias is believed to be an anthocyanin, very closely related to the dye known as "Hoffman's Violet". The coloring matter from the dahlia is green with alkalies and red with acids.

EXPERIMENT 43—Dahlia Test Paper.

Moisten a piece of soft, white paper with a little of the purple dahlia prepared in the preceding experiment. When dry it should be a rich blue violet color. This paper can be used to test for acids or alkalies; being red with acids and green with alkalies.

EXPERIMENT 44—Cherry Indicator.

Obtain some cherries; the dark red sweet ones are best for this experiment. Place a number of cherries in a dish and crush the juice from them. Test a little of this juice with ¼ measure of Sodium Carbonate (No. 4) and then add 1 measure of Sodium Bisulfate (No. 7) and notice the color changes with alkalies and acids.

EXPERIMENT 45—Elderberry Indicator.

Elderberries contain a juice which can be used as an indicator. If you can obtain a few elderberries, crush them and strain the juice. Test this juice with 1 measure of Sodium Carbonate (No. 4) and then with 1 measure of Sodium Bisulfate (No 7). You will find that this indicator is greenish blue with alkalies and red with acids.

EXPERIMENT 46—Elderberry Test Paper.

If you can extract enough of the elderberry juice, put it in a test tube and moisten a piece of soft, white paper with it. When the paper is dry, test a small strip of it with alkalies and acids.

EXPERIMENT 47—Huckleberry Indicator.

If you can get some huckleberries during huckleberry season, put a few in a dish, add a little water and grind the mass until a rich liquid is obtained. Pour off a little of this liquid into a clean test tube and test with sodium carbonate, and then with sodium bisulfate or vinegar. Notice the different colors.

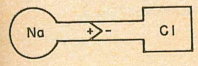
EXPERIMENT 48—Red Cabbage Indicator.

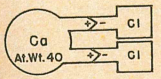
An indicator may be prepared from red cabbage leaves. If you can obtain a red cabbage, break up a few of the leaves into small pieces and soak them overnight in a tumbler half full of water. Stir this liquid until it becomes colored a deep purple. It may be necessary to put the cabbage leaves in hot water in order to extract enough of the color.

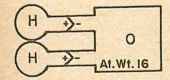
Now pour a little of the clear purple liquid into a test tube and add ¼ measure of Sodium Carbonate (No. 4). You will find that the solution has turned green. Now add 1 measure of Sodium Bisulfate (No. 7), and you will find that the green liquid turns red. This indicator is green in the presence of alkalies and red with acids.

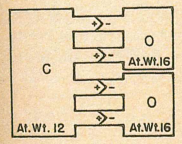
COMBINING CAPACITY OR VALENCE

Atoms of some elements, as we have learned, can lose one or more negative electrons while atoms of other elements can gain various numbers of negative electrons. Perhaps there are some questions which should now be answered. Why should sodium chloride be NaCl but calcium chloride $CaCl_2$? If water is H_2O, why should carbon dioxide be CO_2 and not C_2O? Years of patient experimenting have shown that each element has a definite ability or capacity for combining with other elements. This combining capacity we call valence. A hydrogen atom has one unit of combining capacity, that is, its valence is one. Since one atom of chlorine can combine with one, and only one, atom of hydrogen in forming the compound HCl, the valence of chlorine must also be one. Furthermore, one atom of sodium combines with one atom of chlorine to form NaCl, and so the sodium must likewise have a valence of one.

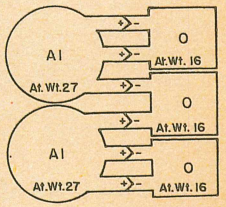




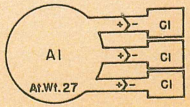




It must not be overlooked, however, that when a sodium atom combines with a chlorine atom, it gives up an electron to the chlorine and hence becomes positively charged, while the chlorine receives a unit of negative electrical charge. We say, therefore, that sodium has a valence of plus one, and that chlorine has a valence of minus one. Hydrogen, like sodium, has ordinarily a valence of plus one, since it, too, tends to give up an electron. Now what about calcium? It has been shown that it can combine with two chlorine atoms by giving up two electrons to them. Its valence then is plus two. Since aluminum can form the compound aluminum chloride, $AlCl_3$, its valence must be plus three. In water there are two atoms of hydrogen for every one of oxygen, and so oxygen must have twice the combining capacity of hydrogen, since each atom of it is able to hold two of hydrogen. The valence of oxygen is thus minus two. How would you write the formula of sodium oxide? Would it be Na_2O? What about calcium oxide? Since calcium has a valence of plus two and oxygen a valence of minus two, each can hold just one atom of the other, forming the compound CaO. Aluminum oxide may be a little harder. If you take two aluminum atoms, each having a valence of plus three, you will have a total positive valence of six. Now how many oxygen atoms must you have in order to get a total negative valence of six? It will take three atoms of oxygen to furnish enough negative combining power to hold the two aluminum atoms. Therefore, the formula of aluminum oxide is Al_2O_3. This compound, colored by impurities, forms the gem minerals, sapphire, ruby, and corundum.



The number of electrons lost by the atoms of some metals may depend upon circumstances. An atom of iron, for example, may lose either two electrons or three electrons. If it loses two and if each is taken up by a chlorine atom, it forms ferrous chloride, $FeCl_2$. If it gives up three, one to each of three



chlorine atoms, ferric chloride, $FeCl_3$, will be formed. Iron, then, is an element with two valences, plus two and plus three. When its valence is two, its compound with oxygen is FeO, ferrous oxide. When its valence is three, it can hold still more oxygen, forming Fe_2O_3, ferric oxide. A number of the elements have two or more valences.

THE DIFFERENCE BETWEEN PHYSICAL CHANGE AND CHEMICAL CHANGE

When water freezes, it is still water. When it evaporates and forms vapor, it is still built up of the same atoms of hydrogen and oxygen. When sugar dissolves in water, it is still sugar. The atoms of carbon, hydrogen, and oxygen remain in the same arrangement in the molecules, and if the water is removed by evaporation the residue is solid sugar. On the other hand, when a wax candle burns, it is no longer wax, and it would be a hard job to "unburn" the candle and get back the original wax. When a piece of iron rusts, new molecules are formed from atoms of iron and atoms of oxygen. Changes like this, which alter the composition of compounds by adding, removing, or rearranging atoms, are called chemical changes. Changes of the first kind, which affect only the physical state of a substance without altering the structure of its molecules, are called physical changes. In more advanced courses in chemistry you will learn that it is occasionally very difficult to decide whether a certain change should be called chemical or physical, but it is of little importance to consider such cases here.

EXPERIMENT 49—A Physical Change.

Put a teaspoonful of sodium chloride, ordinary table salt, in a tumbler half full of water and stir until the salt is all dissolved. Now pour this solution into a small sauce pan and heat it on the stove until the water completely evaporates. When the salt becomes nearly dry, there may be some popping, but this does no harm other than to cause some of the salt to fly around.

Examine carefully the salt that remains in the pan and satisfy yourself that it is the same as the salt with which you started. We learn, therefore, that merely dissolving a substance does not necessarily produce a chemical reaction, even though the physical form of the substances changes completely.

EXPERIMENT 50—A Chemical Change.

Put 2 measures of Azurite (No. 39) (copper carbonate) into a clean dry test tube. Heat this dry powder over your alcohol lamp flame for a few minutes and notice what occurs. First, water is given off as you can see from the moisture formed in the upper part of the test tube. Next, the copper carbonate will gradually become black as it is converted into black copper oxide. In this case, the copper carbonate has been changed chemically, being broken down into copper oxide and carbon dioxide, and you no longer have the original substance with which you started.

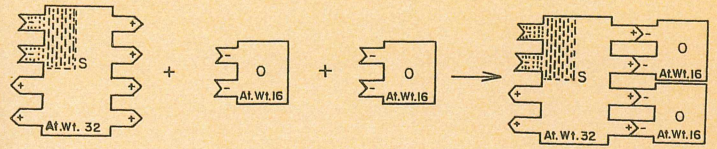
$$\underset{\text{(copper carbonate)}}{CuCO_3} + \text{heat} \rightarrow \underset{\text{(copper oxide)}}{CuO} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

CHEMICAL REACTIONS — COMBINATION OF ELEMENTS

One of the most common chemical changes is the combination of one element with another to form a compound. If a small quantity of sulfur is lighted, it will burn with a bluish flame until the sulfur is all gone. The oxygen of the air has combined with the sulfur forming sulfur dioxide, which is a gas:

$$S + O_2 \rightarrow SO_2\uparrow$$

Since an atom of sulfur in this case has a valence or combining capacity of 4, it requires 2 atoms of oxygen with a valence of 2 to unite with it and the sulfur dioxide molecule which is formed contains 1 atom of sulfur and 2 atoms of oxygen.

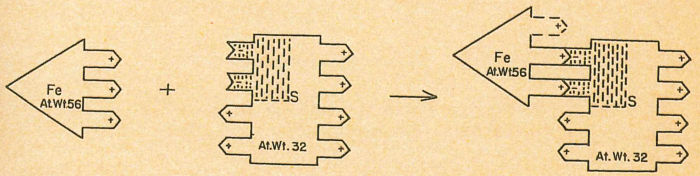


EXPERIMENT 51—The Combination of Iron and Sulfur.

Powdered iron metal, iron filings or steel wool cut up into very small particles will combine with sulfur when heated, forming iron sulfide. Mix together thoroughly on a sheet of paper 2 measures of finely divided iron and 2 measures of Sulfur (No. 1). Heat this mixture in the spoon until the sulfur no longer burns and a glow is noted in the mass. Keep your face away while you are heating. Examine the mass after it cools and you will find it is an entirely different substance from either iron or sulfur. If a small portion of the iron sulfide is put into the test tube with 1 measure of Sodium Bisulfate (No. 7) and a little water and the solution warmed, you will notice the odor of hydrogen sulfide, proving that you have iron sulfide present.

$$\underset{\text{(iron)}}{Fe} + \underset{\text{(sulfur)}}{S} \rightarrow \underset{\text{(iron sulfide)}}{FeS}$$

Notice that in this reaction sulfur has a valence of 2, while iron (ferrous) also has a valence of 2.

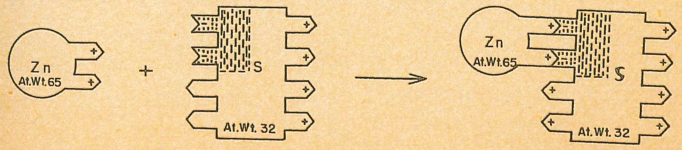


EXPERIMENT 52—Combination of Zinc and Sulfur.

Cut a few very fine slivers from a sheet of zinc. Mix these with an equal quantity of Sulfur (No. 1) and heat the mixture over a flame. The reaction in this case is more noticeable than the sulfur and iron and the product, of course, is zinc sulfide. This may be proven by adding a dilute acid and heating to obtain the odor of hydrogen sulfide as in the preceding experiment.

$$\underset{\text{(zinc)}}{Zn} + \underset{\text{(sulfur)}}{S} \rightarrow \underset{\text{(zinc sulfide)}}{ZnS}$$

Here, also, sulfur has a valence of 2.



EXPERIMENT 53—Combination of Zinc and Oxygen.

Cut a fine sliver of zinc and grip one end of it in your test tube holder. Now, holding the test tube holder in your hand, introduce the other end of the sliver of zinc into a flame. Notice that the zinc burns with a bluish flame or in other words, unites with the oxygen of the air, forming a white powdery substance, which is zinc oxide.

$$\underset{\text{(zinc)}}{2\,Zn} + \underset{\text{(oxygen)}}{O_2} \rightarrow \underset{\text{(zinc oxide)}}{2\,ZnO}$$

EXPERIMENT 54—Combination of Sulfur and Oxygen.

When 1 or 2 measures of sulfur are put in the spoon and heated until the sulfur catches on fire and burns, the sulfur is combining with the oxygen of the air, forming sulfur dioxide, SO_2.

$$\underset{\text{(sulfur)}}{S} + \underset{\text{(oxygen)}}{O_2} \rightarrow \underset{\text{(sulfur dioxide)}}{SO_2\uparrow}$$

CHEMICAL REACTIONS — DECOMPOSITION

We have seen above how 2 elements combine. Now we shall illustrate how a molecule can be split up or decomposed into its elements. By passing an electric current through water the hydrogen can be separated from the oxygen:

$$2\,H_2O \rightarrow 2\,H_2 + O_2$$

In this reaction O_2 is used to designate a molecule of oxygen gas because each molecule of the gas is made up of 2 atoms. It is, therefore, necessary to have 2 molecules of hydrogen, each of which consists of 2 atoms, in order to balance the equation.

EXPERIMENT 55—Decomposition of Sodium Thiosulfate.

Place 2 measures of Sodium Thiosulfate (No. 16) in a test tube. Use the test tube holder so as not to burn your fingers, and heat the bottom of the tube over a flame. Notice that the solid first becomes moist and steam is given off. This is because sodium thiosulfate contains water of hydration which is driven off when the substance is heated.

Continue heating for a few minutes longer and you will notice that a yellowish powder, sulfur, is being deposited on the upper part of the test tube. Remove the tube from the flame and smell the gas which is given off at the mouth of the tube. You will recognize an odor like rotten eggs due to the gas, hydrogen sulfide.

You learn from this experiment that by heating sodium thiosulfate, it is broken up into several different substances. First — water, which was driven off as steam. Second — sulfur, which was deposited around the upper part of the test tube. Third — hydrogen sulfide gas, which you could smell at the mouth of the tube. Fourth — a reddish mass in the bottom of the tube which consists of a mixture of sodium sulfate and sodium sulfide.

EXPERIMENT 56—The Decomposition of Sugar.

Place 2 measures of powdered or granulated sugar in a clean dry test tube. Heat the tube over a flame, keeping your face at a little distance, and watch carefully the changes which take place in the sugar. First — steam is given off, showing that the sugar contains water. Next the sugar gradually becomes darker and finally you have remaining in the test tube a black substance which is carbon, an element. Thus you have learned that sugar is composed of water (hydrogen and oxygen) and carbon.

$$\underset{\text{(sugar)}}{C_{12}H_{22}O_{11}} + \text{heat} \rightarrow \underset{\text{(carbon)}}{12\,C} + \underset{\text{(water)}}{11\,H_2O}$$

EXPERIMENT 57—Decomposition of Calcium Carbonate.

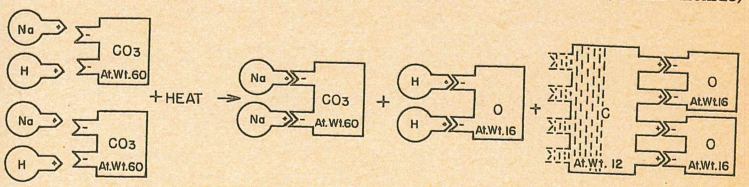
Calcium carbonate is the main constituent of limestone and oyster shells and can be decomposed by heat to form lime and carbon dioxide. Obtain a small piece of limestone or oyster shell and break it with a hammer so that you have a number of very small slivers. Pick out one of these slivers which has a very thin knife-like edge or point. Hold the sliver with a pair of pliers and heat the thin edge or point in your alcohol lamp flame for several minutes, keeping it red hot during this time. If you have a gas flame available, this experiment can be done more quickly with a hotter flame. After the sliver of limestone or oyster shell has cooled, examine carefully the edge which you heated and you will notice that this particular part has become white and chalky and that some white powder can be scraped off with a knife. The heat has driven off the carbon dioxide from the calcium carbonate, leaving calcium oxide, commonly known as lime.

$$\underset{\text{(calcium carbonate)}}{CaCO_3} + \text{heat} \rightarrow \underset{\text{(calcium oxide)}}{CaO} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

EXPERIMENT 58—The Decomposition of Sodium Bicarbonate.

Put 2 measures of sodium bicarbonate (baking soda) in a test tube and heat over a flame. You will notice that moisture is given off although the baking soda with which you started was perfectly dry. As you heat the sodium bicarbonate, light a match and thrust it into the mouth of the test tube. It will immediately go out, showing that carbon dioxide gas is being given off. After a few minutes a white powdery mass will remain in the test tube, which consists of sodium carbonate.

$$\underset{\text{(sodium bicarbonate)}}{2\,NaHCO_3} + \text{heat} \rightarrow \underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$



To prove that you have decomposed sodium bicarbonate, dissolve 1 measure of baking soda in a clean test tube half full of water. Add 1 drop of Phenolphthalein Solution (No. 22) and notice that you get only a very faint pink color, showing that sodium bicarbonate is almost neutral. Now add a little water to the test tube containing the residue remaining after you heated the sodium bicarbonate. Make sure this test tube is cool before you add water. When this substance has dissolved, add 1 drop of phenolphthalein solution and note that you get an intense red color, showing that you have a strongly alkaline salt which corresponds with sodium carbonate.

CHEMICAL REACTIONS — DISPLACEMENT

When a metal is acted upon by an acid, the hydrogen of the acid is displaced by the metal and the hydrogen is liberated as a gas:

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$$

The metal in the above reaction is zinc which displaces the hydrogen of the sulfuric acid forming zinc sulfate and liberating hydrogen. Zinc has a valence or combining power of 2 and hydrogen has a valence or combining power of 1, therefore, 1 atom of zinc replaces 2 atoms of hydrogen in the sulfuric acid molecule H_2SO_4. This then becomes zinc sulfate, $ZnSO_4$, and the 2 atoms of hydrogen are liberated in the form of a hydrogen molecule.

EXPERIMENT 59—Displacement of Hydrogen by Zinc.

Dissolve 2 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Now cut several fine slivers of zinc metal, using an old pair of scissors or tin snips. Drop these slivers of zinc into the test tube containing sodium bisulfate solution and warm the solution over a flame. You will notice bubbles of gas beginning to come from the pieces of zinc. This is hydrogen being displaced from the sodium bisulfate while the zinc is going into solution in the form of zinc sulfate. After this action has been going on for a few moments, hold a lighted match at the mouth of the test tube. If the hydrogen gas is coming off fast enough, you will notice a slight pop as the hydrogen is ignited.

$$\underset{\text{(sodium bisulfate)}}{2\ NaHSO_4} + \underset{\text{(zinc)}}{Zn} \rightarrow \underset{\text{(sodium sulfate)}}{Na_2SO_4} + \underset{\text{(zinc sulfate)}}{ZnSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 60—Displacement of Hydrogen by Iron.

Dissolve 2 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water and add a few iron filings or very small pieces of steel wool. Warm this solution over a flame and you will notice bubbles of gas forming on the iron. In this case the action is not as vigorous as it was with the zinc because iron is a less active metal than zinc.

$$\underset{\text{(sodium bisulfate)}}{2\ NaHSO_4} + \underset{\text{(iron)}}{Fe} \rightarrow \underset{\text{(sodium sulfate)}}{Na_2SO_4} + \underset{\text{(iron sulfate)}}{FeSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 61—Displacement of Copper by Iron.

Put 1 measure of Azurite (No. 39) in a test tube and add 3 measures of Sodium Bisulfate (No. 7). Fill the test tube half full of water, close the mouth with your thumb and shake until the solids are entirely dissolved.

Azurite is a mineral containing a large percentage of copper carbonate. This substance while insoluble in water will dissolve readily in sodium bisulfate solution, forming a solution of copper sulfate (blue vitriol).

Now add 1 measure of powdered iron metal, or small pieces of steel wool, to the solution of copper sulfate which you have just made and shake the tube for a few minutes. The blue color of the copper sulfate solution will gradually disappear and the iron will become covered with a red deposit of copper.

Iron compounds, when dissolved in water, give a blue color with Sodium Ferrocyanide (No. 6). Add ½ measure of this substance to the test tube and shake it. This will prove that some of the iron is now dissolved in the liquid.

$$\underset{\text{(copper sulfate)}}{CuSO_4} + \underset{\text{(iron)}}{Fe} \rightarrow \underset{\text{(ferrous sulfate)}}{FeSO_4} + \underset{\text{(copper)}}{Cu}$$

From the above equation you can easily see what happened. The iron simply pushed the copper out of the copper sulfate, forming ferrous (iron) sulfate and setting free the copper in the metallic state.

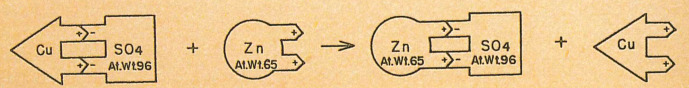
EXPERIMENT 62—Displacement of Copper by Zinc.

Prepare some copper sulfate solution by adding 3 measures of Sodium Bisulfate (No. 7) and 1 measure of Azurite (No. 39) to a test tube half full of water as in the preceding experiment.

Now add a small piece of zinc about the size of the end of the measure to the copper sulfate solution, close the mouth of the test tube with your thumb and shake it, then let stand for a few minutes. This time you will find that the zinc has become covered with a dark brown deposit of copper and the blue of the copper sulfate solution has disappeared.

$$\underset{\text{(copper sulfate)}}{CuSO_4} + \underset{\text{(zinc)}}{Zn} \rightarrow \underset{\text{(zinc sulfate)}}{ZnSO_4} + \underset{\text{(copper)}}{Cu}$$

This reaction, you see, is exactly like the preceding one except that it is zinc which displaces copper, forming, of course, zinc sulfate instead of ferrous sulfate.



CHEMICAL REACTIONS — DOUBLE DECOMPOSITION

Suppose some sodium carbonate (washing soda) is dissolved in water and added to a solution of calcium chloride. Immediately the liquid becomes milky with fine particles of a solid compound, calcium carbonate (precipitated chalk). If this solid is separated from the liquid by filtering, or by allowing the precipitate to settle, and the water is evaporated from the remaining clear liquid, crystals of sodium chloride (table salt) will be left in the vessel. Here is what happened, expressed in the form of a chemical equation:

$$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2\ NaCl$$

The calcium and the sodium seemed to have exchanged partners during the reaction.

This equation we have just written is good enough for most purposes, but it does not tell quite the whole story, because compounds such as calcium chloride and sodium carbonate are made up of ions. Each calcium atom has a strong tendency to part with two of its electrons, and each chlorine atom has an equally strong tendency to pick up one electron. It is easy to see, then, how two chlorine atoms might each take one of the electrons which a calcium atom is anxious to get rid of. This forms two negatively charged chloride ions and one positively charged calcium ion. The positive and negative charges attract each other, and in a crystal of solid calcium chloride the ions hold each other tightly together. When the crystal is dissolved in water, the ions are permitted to separate, and so there are calcium and chloride ions wandering around through the liquid.

In the same way, when sodium carbonate is dissolved, the solution contains sodium ions, Na^+, and carbonate ions, CO_3^{--}. Notice that the carbon atom and three oxygen atoms of the carbonate group hang together in a single ion and go through the whole reaction without breaking apart.

What really happens in this experiment, then, is the mixing of sodium ions and carbonate ions with calcium ions and chloride ions. The calcium ions immediately combine with the carbonate ions to form the insoluble compound, calcium carbonate:

$$2\ Na^+ + CO_3^{--} + Ca^{++} + 2\ Cl^- \rightarrow CaCO_3 + 2\ Na^+ + 2\ Cl^-$$

The sodium and chloride ions simply remain in the solution, and when the water is evaporated they come together to form crystals of solid sodium chloride.

A great many of the common reactions which go on in solution are really reactions between ions, but frequently one does not bother to show the ions in the equations.

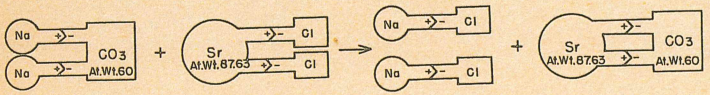
EXPERIMENT 63—Precipitation of Strontium Carbonate from Sodium Carbonate and Strontium Chloride.

Put 1 measure of Sodium Carbonate (No. 4) in a test tube. Fill the tube ¼ full of water and shake it until the solid is dissolved. In another test tube put 1 measure of Strontium Chloride (No. 10). Fill the tube ¼ full of water and shake the tube to dissolve the solid.

You now have a solution of two different salts. Pour the contents of one test tube into the other and notice that a white solid or precipitate is formed. The sodium carbonate and strontium chloride react forming sodium chloride and strontium

carbonate. The sodium chloride was soluble in water and remained dissolved, but the strontium carbonate was insoluble in water and so it separated from the solution as a solid or precipitate.

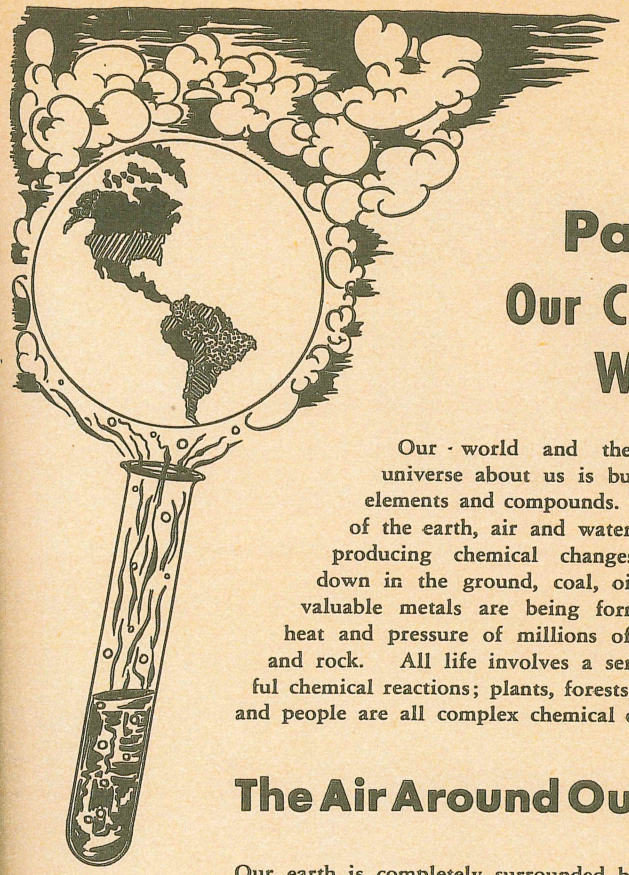
$$\underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(strontium chloride)}}{SrCl_2} \rightarrow \underset{\text{(sodium chloride)}}{2\ NaCl} + \underset{\text{(strontium carbonate)}}{SrCO_3}$$



EXPERIMENT 64—Precipitation of Strontium Sulfate from Aluminum Sulfate and Strontium Chloride.

Dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube ¼ full of water. In another test tube ¼ full of water dissolve 1 measure of Strontium Chloride (No. 10). You now have two colorless solutions. Pour the contents of one tube into the other and you immediately obtain a heavy white precipitate, showing that a chemical reaction has taken place. In this reaction you have formed strontium sulfate, which is insoluble in water and appears as a heavy white precipitate, while aluminum chloride remains in the solution.

$$\underset{\text{(aluminum sulfate)}}{Al_2(SO_4)_3} + \underset{\text{(strontium chloride)}}{3\ SrCl_2} \rightarrow \underset{\text{(strontium sulfate)}}{3\ SrSO_4} + \underset{\text{(aluminum chloride)}}{2\ AlCl_3}$$



Part II
Our Chemical World

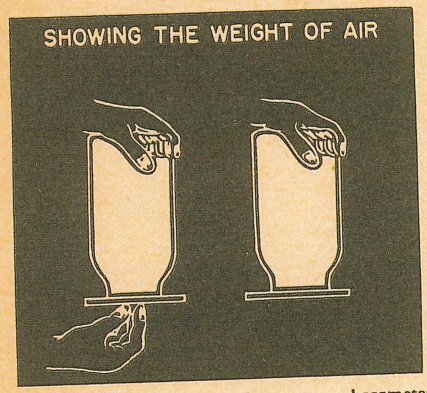
Our world and the whole great universe about us is built of chemical elements and compounds. On the surface of the earth, air and water are constantly producing chemical changes, while deep down in the ground, coal, oil, and ores of valuable metals are being formed under the heat and pressure of millions of tons of earth and rock. All life involves a series of wonderful chemical reactions; plants, forests, birds, animals and people are all complex chemical engines.

The Air Around Our Earth

Our earth is completely surrounded by a blanket or layer of gases. These gases are what we call air. Everyone realizes the importance of the air. It makes possible both animal and plant life; it keeps the sun from scorching us during the day and it reduces the earth's loss of heat into empty space during the night. This blanket of air is not very thick, as it extends upward from the earth's surface only about 50 miles. The weight of the air, however, over the whole surface of the earth is enormous. You may have heard that the atmosphere exerts a pressure of nearly 15 pounds per square inch at the surface of the earth. In other words, a column of air, one inch square and as high as the atmosphere extends, weighs almost 15 pounds. Thus the air directly above an acre of ground (an area about the size of the average baseball grounds) weighs approximately 90,000,000 pounds.

EXPERIMENT 65—Showing the Weight of Air.

Obtain a milk bottle or drinking glass with a perfectly smooth edge. Working at a sink or wash basin, fill this container with water so that the water is even with, or stands just a tiny bit higher than, the rim of the glass or bottle. Now, cut a small piece of cardboard large enough to cover the mouth of the glass or bottle and carefully slide this across the top so that it forms perfect contact with the water inside of the container. Place the finger of one hand lightly on the cardboard and with the other hand turn the glass or bottle upside down. Now carefully remove your finger and if you have followed directions correctly, the water will not spill from the container. This is because the weight of the air presses against the cardboard strongly enough to prevent the water falling out of the glass or bottle.



The weight or pressure of the air is strong enough to support a column of mercury 30 inches high. This is the principle upon which the mercury barometers for measuring the air pressure work.

EXPERIMENT 66—The Weight of the Air Crushes a Metal Can.

If you can obtain a large metal can with a tight screw cap, you can illustrate the weight of the air very easily. The can must have a capacity of at least a gallon and of course must be tight, and also have a screw cap which fits well, or a tight-fitting wooden plug or rubber stopper. (A gallon size square oil can is ideal.)

Put enough water in the can to form a layer about an inch deep in the bottom. Now set the can on a stove and heat it until the water inside is boiling briskly and steam is coming out of the opening on top. At this point, turn off the heat or remove the can quickly from the burner and immediately screw the cap on tightly. Now, holding the can with a cloth so as not to burn yourself, hold it under a cold water faucet and allow the cold water to quickly cool the can. This condenses the steam inside the can which had largely replaced the air and as the steam is condensed, leaving a partial vacuum inside of the can, the weight of the air on the outside will cause the can to collapse or curve inward. If you use a large can, it may completely collapse, but a small can will probably show only slight dents in the side.

OXYGEN

What is this important mixture of gases that we call air? It consists principally of two gases, nitrogen and oxygen. About one-fifth of air is oxygen and four-fifths of air is nitrogen. There are several other gases present in small amounts, including argon, neon, krypton, and of course, carbon dioxide and water vapor.

It is the oxygen in the air which supports life. When we breathe, the oxygen of the air taken into the lungs is absorbed by the blood and carried to all parts of the body. The various tissues slowly unite with the oxygen, forming carbon dioxide. Heat is liberated by this reaction. The carbon dioxide is expelled in our breath. If the air should be deprived of oxygen, we could not live, but on the other hand, if there were no nitrogen in the air and we were forced to breathe pure oxygen, our tissues would be consumed so rapidly that we would soon expend our energy.

Plants absorb carbon dioxide from the air through their leaves, and combine it with water to form, by a series of complicated reactions, various compounds from which their tissues are very largely built up. In this process oxygen is liberated and returned to the air. Thus, plants and animals mutually assist in keeping the quantity of oxygen and of carbon dioxide in the air constant.

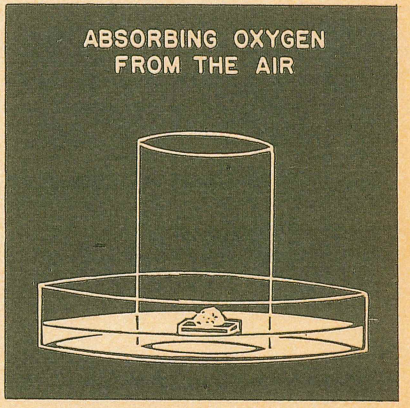
Of all the elements, oxygen is the most abundant and the most widely distributed. This element, in the free state, is a colorless, odorless gas. Air, as we have learned, is one-fifth free oxygen, and water, which we will soon take up, is eight-ninths combined oxygen. Common materials like sandstone, limestone, brick and mortar contain about 50% oxygen. It has been estimated that about one-half of our world is oxygen.

Oxygen is a very active element, combining readily with many other elements. These reactions are called oxidation. Burning or ordinary combustion is a process of oxidation. Thus, the burning of coal is mainly the oxidation of carbon.

It can readily be seen that the activity of oxygen affects our lives beneficially in many ways and yet, because of its abundance in the free state in the air and its high chemical activity, oxygen also creates grave problems. Uncontrolled combustion, such as fires that destroy our buildings and sweep our forests, and many of the disastrous explosions that wreck mines and buildings, sometimes causing great loss of life, are examples of the uncontrolled reactions of oxygen. The corrosion of metals, as, for example, the "rusting" of iron, really is a reaction of oxidation, or very slow combustion.

EXPERIMENT 67—Absorbing Oxygen from the Air.

Obtain a small flat piece of wood or cork which will fit loosely inside an ordinary tumbler. Fill a saucer or small dish with water, float the piece of wood on the water and on the wood, place a small wad of steel wool about the size of the end of your thumb. Make sure that it is steel wool (iron) and not some other metal such as copper. You can use 4 measures of powdered iron or iron filings if more convenient. Moisten the steel wool with a few drops of water. Now invert a tumbler over the floating wood in order to trap the air in the tumbler and then set the saucer and tumbler and floating wood in a quiet place for a few days.



After 2 or 3 days, when the steel wool has begun to rust, notice that the water is slowly rising inside of the glass, showing that the volume of the enclosed air is becoming less. This is because some of the oxygen in the air that was trapped inside of the glass has united with the iron of the steel wool, forming iron oxide or rust.

EXPERIMENT 68—Measuring the Amount of Oxygen in the Air.

You can easily determine the amount of oxygen in ordinary air by the following procedure:

Set up the apparatus and material that you used in the last experiment and before setting it aside, measure the distance from the surface of the water to the inside top of the glass. Record this distance and then set the glass aside as it will take some time for the iron to absorb all of the oxygen.

After several days, again measure the distance from the surface of the water to the top of the glass. To find the percentage of oxygen, subtract this value from the original reading, divide by the original reading and then multiply the result by 100. This should give you a value of approximately 20% for the amount of oxygen in air. The exact value cannot be determined because no corrections have been made for changes in temperature and pressure which take place during the experiment and which cause changes in the volume of the enclosed air.

EXPERIMENT 69—Removing Oxygen from Air with a Candle Flame.

Set a candle in the middle of an ordinary pan, first allowing a few drops of melted wax to fall on the pan to stick the candle tight. Next pour about 2 inches of water in the pan. Now light the candle and set over it an inverted quart milk bottle or fruit jar. The bottle or jar must be high enough so that the bottom will not come too close to the candle flame.

Notice that almost at once the candle flame grows dim and that after a few minutes it goes out, the oxygen of the air inside the bottle having been consumed. Notice also that after the flame has gone out the water begins to rise inside the bottle and finally stands at a considerably higher level in the bottle than in the pan. This is due to the removal of the oxygen from the air inside of the bottle, leaving a partial vacuum so that the air pressure on the outside pushes some of the water from the pan up into the bottle. In the flame, the oxygen of the air unites with the candle wax, which contains carbon, thus forming carbon dioxide, a gas which will dissolve in water, and steam which condenses into water.

EXPERIMENT 70—How Leaves Give Off Oxygen.

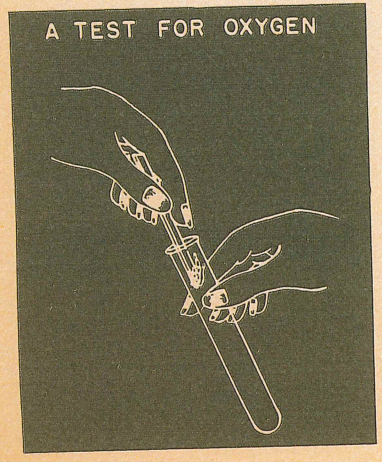
Obtain a quart milk bottle or jar and put a bunch of green leaves in it. Fill the jar full of water, placing a small piece of cardboard over the mouth and carefully turn it upside down in a pan of water so that the mouth of the jar is under the water in the pan. Carefully remove the cardboard. As long as you keep the mouth of the jar below the level of the water in the pan, no air will get into the jar or bottle. Place the pan with the bottle or jar and leaves in the sunlight and notice that a gas is given off which collects above the water in the bottle or jar. This gas is oxygen.

Try this experiment again, but keep the jar or bottle in the dark and you will find that no oxygen is given off. In this case, some carbon dioxide is given off by the leaves, but this dissolves in the water and does not collect at the top of the bottle.

EXPERIMENT 71—Making Oxygen by a Chemical Reaction.

If you can obtain some hydrogen peroxide, which is sometimes found in the medicine cabinet, you can perform this very interesting experiment. Fill a test tube ¼ full of hydrogen peroxide. In another test tube ¼ full of water, dissolve 2 measures of Sodium Carbonate (No. 4) and add 4 or 5 drops of Cobalt Chloride Solution (No. 27). Shake the tube vigorously. Pour a few drops of this pinkish liquid into the hydrogen peroxide in the other test tube. It is best to hold the test tube over a dish or over the sink while you mix the liquids because they foam up and may spill over the top. The gas being driven off is oxygen. This oxygen is obtained from the decomposition of the hydrogen peroxide. The cobalt chloride and the sodium carbonate together form basic cobalt carbonate, which helps to decompose the hydrogen peroxide without being itself consumed in the reaction. A substance which assists a reaction in this way is called a "catalyst".

$$\underset{\text{(hydrogen peroxide)}}{2\,H_2O_2} \rightarrow \underset{\text{(water)}}{2\,H_2O} + \underset{\text{(oxygen)}}{O_2\uparrow}$$



EXPERIMENT 72—A Test for Oxygen

Obtain a wooden splint about 6 inches long, ¼ inch wide and very thin. In the preceding experiment, while oxygen is being generated in the test tube, light the wooden splint at one end and after a few seconds blow it out so that a glowing spark remains on the end. Quickly thrust this glowing spark into the test tube in which oxygen is being generated and the spark will glow much brighter or perhaps burst into flame. This is the usual test for oxygen.

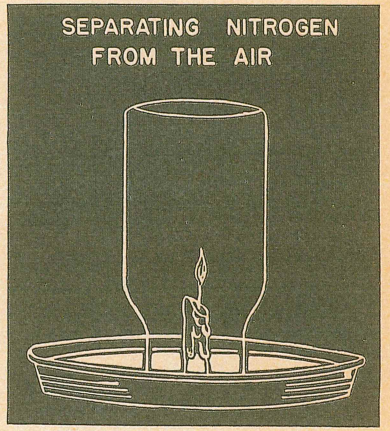
NITROGEN

Nitrogen is a colorless, odorless gas, which, as you have already learned, makes up nearly four-fifths of the air. This element is of great importance because it is needed for the growth of plants and it also has many uses in

industry. Nitrogen is so inactive, however, that it tends to stay unchanged in the atmosphere, and even though the world is surrounded with an unlimited supply of nitrogen, it is impossible, except on a very limited scale, to make use of this element in many important ways. Early in this century, a chemist named Fritz Haber invented a process in which nitrogen of the air is made to combine with hydrogen to form ammonia. The Haber process has given the world an unlimited supply of nitrogen compounds at low cost, thus assuring ample fertilizer as well as all the nitrogen compounds we need for industrial purposes.

EXPERIMENT 73—Separating Nitrogen from Air.

Repeat the former experiment, "Removing Oxygen from Air with Candle Flame". The gas remaining in the bottle or jar is mostly nitrogen. Place a small glass plate or piece of cardboard over the mouth of the bottle while it is still under water, and then turn the bottle right side up, keeping the cover in place so that the gas cannot escape.



EXPERIMENT 74—Wood Will not Burn in Nitrogen

Obtain a thin splinter of wood about 6 or 8 inches long and light one end of it. Remove the cover from the bottle containing nitrogen separated in the preceding experiment and thrust the burning splinter into the bottle. The flame will go out almost immediately, showing that nitrogen does not support combustion.

COMPOUNDS OF NITROGEN

The compounds of nitrogen are among the most interesting and valuable known to the chemist. With hydrogen and oxygen, nitrogen forms nitric acid. Since nitrogen is essential to the growth of plants, we find compounds of nitrogen present in almost all fertilizers. The products of animal life also contain large amounts of nitrogen. As we have learned in the Haber process, the nitrogen of the air is made to combine with hydrogen to form the compound ammonia. Ammonia is extensively used as a refrigerant, in making artificial ice, and in cold storage plants. It is used in the production of nitric acid and many other important compounds containing nitrogen.

EXPERIMENT 75—Making Ammonia.

Place 2 measures of Ammonium Chloride (No. 9) and 2 measures of Calcium Oxide (No. 20) in a test tube. Notice that neither of these chemicals has any odor before they are mixed together. Gently heat the mixture over a candle flame and after a moment, smell cautiously. Do you recognize the odor of ammonia? Moisten a piece of red Litmus Paper (No. 18) and hold this in the mouth of the test tube in which ammonia is being made. Does the red litmus paper turn blue? This simple experiment illustrates one of the ways in which ammonia is often made in chemical laboratories.

$$\underset{\text{(ammonium chloride)}}{2\,NH_4Cl} + \underset{\text{(calcium oxide)}}{CaO} + \text{heat} \rightarrow \underset{\text{(calcium chloride)}}{CaCl_2} + \underset{\text{(ammonia)}}{2\,NH_3\uparrow} + \underset{\text{(water)}}{H_2O}$$

EXPERIMENT 76—Another Method of Making Ammonia.

Place 2 measures of Ammonium Chloride (No. 9) in a test tube as before, but this time add 2 measures of Sodium Carbonate (No. 4). Heat the mixture gently, remove the test tube from the flame and after a few seconds notice the smell of ammonia at the mouth of the test tube.

EXPERIMENT 77—Making Ammonia from Organic Matter.

Obtain some small bits of horn (trimmings from the fingernails,) or a small quantity of hair. Place this in a test tube together with 2 measures of Calcium Oxide (No. 20) and add 2 or 3 drops of water. Heat this mixture gently and see if you can detect the odor of ammonia at the mouth of the test tube after removing the test tube from the flame. Moisten a small strip of red Litmus Paper (No. 18) and hold this at the mouth of the tube. If ammonia is present, the red litmus paper will be turned blue, but the color will fade away. Most organic matter contains nitrogen and when heated with calcium oxide, ammonia is formed.

EXPERIMENT 78—Making Ammonia in Your Hand.

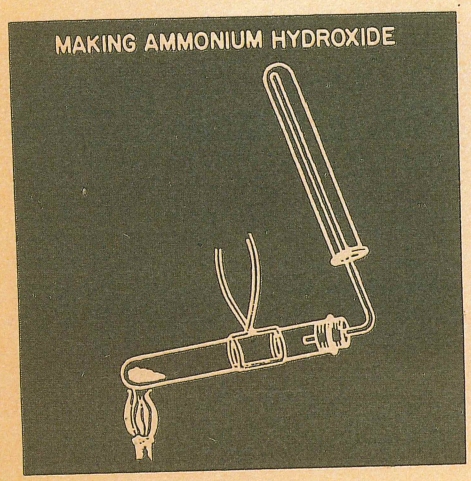
Put 1 measure of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9) in the palm of your hand and mix the substances together with your finger. Smell the mixture occasionally and notice that ammonia gas is being given off. Be sure to wash your hands immediately after performing this experiment.

EXPERIMENT 79—The Solubility of Ammonia.

Ammonia gas is very soluble in water and the liquid or household ammonia, which is commonly used, is a solution of this gas in water. Place in a test tube 2 measures of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9). Heat this mixture gently, keeping the test tube partly closed with your thumb. After you smell ammonia gas at the mouth of the tube, remove it from the flame and close the tube tightly with your thumb. Turn the tube upside down and hold the mouth under the surface of a pan of water. Now remove your thumb from the mouth of the tube and you will find the water will rise part way up inside the tube due to the ammonia dissolving in the water.

EXPERIMENT 80—Making Ammonium Hydroxide.

Put 2 measures of Calcium Oxide (No. 20) and 2 measures of Ammonium Chloride (No. 9) in a test tube. Fit the test tube with the gas delivery tube and stopper.



Hold the test tube with a test tube holder and heat it over a flame with the delivery tube pointed upward. When you can smell ammonia gas coming from the end of the tube, place over it a dry test tube. Since ammonia is much lighter than air, it will rise into this second test tube and push down the air which was originally in it. Continue heating until the odor of ammonia is strong around the mouth of the inverted test tube. Then remove the test tube from the delivery tube, keeping it upside down. As quickly as possible place your thumb over the test tube mouth to prevent the ammonia from escaping. Turn it right side up, still covered by your thumb. Then remove your thumb just a little and fill the test tube ⅛ full of water. Replace your thumb and shake. You will feel a suction on your thumb as the gas dissolves in the water, forming ammonium hydroxide in accordance with the following equation:

$$\underset{\text{(ammonia)}}{NH_3} + \underset{\text{(water)}}{H_2O} \rightleftarrows \underset{\text{(ammonium hydroxide)}}{NH_4OH}$$

The reaction in this case is written with a double arrow to show that it is reversible, that is, ammonium hydroxide can be decomposed again to give off ammonia gas. Drop a small piece of red Litmus Paper (No. 18) into the ammonium hydroxide which you have just made. Does the paper turn blue?

EXPERIMENT 81—Solvent Action of Ammonium Salts.

If you have some Epsom salts (magnesium sulfate) in your medicine cabinet, dissolve 1 measure of this compound in a test tube ¼ full of water and add a few drops of household ammonia, or use the ammonium hydroxide which you prepared in the preceding experiment. Shake the test tube a little and the liquid will become cloudy with a suspension of magnesium hydroxide. Now dissolve 2 measures of Ammonium Chloride (No. 9) in a test tube ¼ full of water and add this solution to the magnesium hydroxide. You will find upon shaking the test tube that the liquid becomes clearer for the ammonium chloride dissolves the precipitated magnesium hydroxide. The solvent action of ammonium salts upon many hydroxide precipitates is important in analytical chemistry.

Ammonium salts are quite volatile and some of them pass into the form of vapor without first melting.

EXPERIMENT 82—Volatilization of Ammonium Chloride.

Put 1 measure of Ammonium Chloride (No. 9) into a clean, dry test tube. Heat gradually over a candle flame, being careful not to get the bottom of the test tube sooty. Notice that the ammonium chloride gradually disappears or turns into a vapor and part of this vapor condenses again on the cool upper part of the test tube.

EXPERIMENT 83—Cobalt Ammonia Green.

Fill a test tube ¼ full of water and add 1 measure of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9). Warm the tube over a flame for a few minutes until the odor of ammonia gas is noticeable at the mouth of the test tube. Be sure to remove the tube from the flame before bringing your face near to smell the gas. Fill another clean test tube ¼ full of water and add 3 or 4 drops of Cobalt Chloride Solution (No. 27). Now pour in just a few drops of the solution containing ammonia and you should obtain a very pretty bluish green color. This is one of the complex cobalt ammonia compounds.

EXPERIMENT 84—Copper Ammonia Blue.

Put 1 measure of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9) in a test tube. Fill the tube ¼ full of water and warm gently over a flame until you notice a slight odor of ammonia at the mouth of the tube. In another test tube, put ½ measure of Azurite (No. 39) and fill the tube ¼ full of water. To this add a few drops of the ammonia solution from the first test tube. Close the mouth of the tube and shake so that the azurite mixes thoroughly with the liquid. Now allow this tube to stand for a few moments and as the excess azurite settles, you will notice that the clear liquid above it is a deep blue color. This color results from a soluble copper ammonia compound which is generally referred to as copper ammonia blue.

The Water On Our Earth

Two elements, which are gases at ordinary temperatures, combine to form a compound which is a liquid. These gases are hydrogen and oxygen and the compound which they form is water. Two atoms of hydrogen unite with one atom of oxygen to form this compound and the chemical formula is H_2O.

Water is the most abundant and also the most important liquid in the world. As you no doubt know, over two-thirds of the surface of our earth is covered with liquid water or with ice. The air which surrounds our earth contains a large amount of water vapor. Sometimes, when the humidity is high, and the air is warm, there may be as much as 4% by volume of water vapor in the air; and at other times, when it is quite dry, there may be almost no water vapor in the air. A man who weighs 140 pounds

contains about 100 pounds of water. Our foods contain large amounts of water. Bread, for example, is almost 35% water and tomatoes will run as high as 94% water.

Water is necessary to all forms of plant and animal life. In the case of plants, the nutrient or plant food elements in the soil must be dissolved in water before they can be carried to the tissues of the plants where they help build additional plant cells. In our bodies, water acts as a solvent for our foods and also aids in the elimination of waste materials, helps to regulate body temperatures and acts as a lubricant.

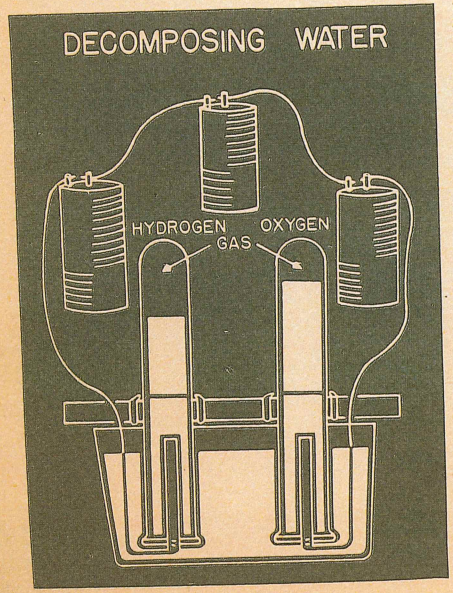
In industry, water enters into many chemical reactions, but its most important use is as a solvent or carrier. Substances, when dissolved, are brought into much more intimate contact and react more readily than in the solid form.

EXPERIMENT 85—Producing Water from Hydrogen and Oxygen.

Fasten a candle in an upright position on a tin pan or other piece of metal, light it and hold over the candle, mouth downward, a cold, dry glass, such as a jelly glass or tumbler. Hold the mouth of the glass just above the flame. After a few seconds, notice the mist of tiny drops of water which is formed on the inside of the glass. Candle wax contains hydrogen and carbon combined in compounds called hydrocarbons. When the candle burns, the hydrogen of the hydrocarbons combines with the oxygen in the air, forming water which, due to the heat of the flame, is in the form of steam and is therefore invisible. To see this water, it must be condensed on a cold, dry object, such as the tumbler which you used.

EXPERIMENT 86—Decomposing Water into Hydrogen and Oxygen.

To perform this experiment you will need some materials that are not included in your Chemcraft outfit. Obtain 2 or 3 good dry cell batteries, 3 or 4 feet of (small) insulated copper wire, and a shallow glass or porcelain dish. Remove the carbon rods from 2 old flashlight batteries. To do this, grasp the ends with a pair of pliers, turn the carbon to break the sealing wax, and then pull the rod straight out. Fill the shallow dish about ½ full of water. Add about 5 measures of Sodium Bisulfate (No. 7) for each glass of water used. Fill 2 test tubes with this solution, close them with your thumb and then turn them upside down in the dish, being careful not to let any air get into them. As long as the mouths of the tubes are kept beneath the surface of the solution they will remain filled with the liquid.



Connect the three dry cells in "series," that is, connect the carbon binding post of one dry cell to the zinc binding post of the next dry cell. Attach long pieces of wire to the free zinc and carbon binding posts of the battery series. Fasten the free ends of these wires to the carbon rods which you have removed from the flashlight batteries, first scraping the insulation from several inches of the wire that is wrapped around the carbon rods. When this has been done, put one of the carbons under each of the test tubes, being very careful not to let any air get into either of the tubes. You will notice almost immediately that bubbles of gas form around the pieces of carbon and rise to the tops of the test tubes, displacing the water downward. In case the bubbles form very slowly or not at all, it will be necessary to connect more dry cells in the circuit in order to increase the flow of current.

Watch the gases as they collect in the tops of the test tubes. You will find that gas collects just twice as fast in one test tube as it does in the other. The water is being decomposed into hydrogen and oxygen. Water contains two atoms of hydrogen to every atom of oxygen. Hence when water is decomposed just twice as great a volume of hydrogen is given off as oxygen. The sodium bisulfate was added to make the water conduct the electric current more readily.

EXPERIMENT 87—Testing the Gases Obtained from Water.

After you have collected a considerable amount of gas in the test tubes of the preceding experiment, remove the carbons from the test tubes, being very careful to keep the mouths of the tubes under water so that no air can enter. Put your thumb over the mouth of the tube containing hydrogen and remove it from the water. Turn it right side up, bring its mouth up to a flame, and remove your thumb. As the hydrogen takes fire, you will see a pale blue flash at the mouth of the tube, accompanied by a sharp pop. This is the characteristic behavior of hydrogen when ignited in the presence of air.

In a similar way remove the tube of oxygen from the pan of water. Have ready a burning splinter several inches long, and when you are ready to remove your thumb from the mouth of the tube, blow out the flame so that only a glowing spark remains. Plunge this spark into the oxygen gas, and the splinter will burst into flame. This is the simplest test for oxygen.

THE PHYSICAL PROPERTIES OF WATER

Depending upon the temperature, water exists in three physical forms. It is a liquid at ordinary temperatures, when cooled to 32° F. it becomes a solid (ice), while when heated to 212° F. it boils and passes into the form of vapor or gas (steam).

When water boils, its vapor tension or pressure is equal to atmospheric pressure and the liquid tends to vaporize completely. It is interesting to notice that water boils more easily at high altitudes than at places near the sea level. This is because the atmospheric pressure is less at high altitudes. Since water boils at a lower temperature at high altitudes, it takes longer to cook foods than at places near sea level, because water can never be heated above its boiling temperature unless it is under pressure. Everyone is familiar with the melting of ice to form water and the boiling of water to form steam. It is interesting to remember to what a large extent our present state of civilization depends upon these physical properties of water.

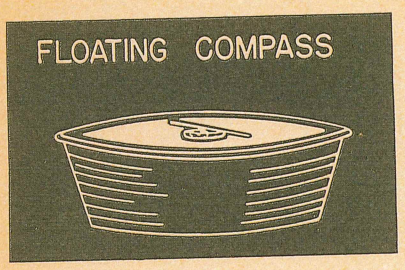
SURFACE TENSION

All liquids behave more or less as if their surface were covered with an elastic film. This arises from a force known as surface tension, which is due to the attraction between the molecules of the liquid. In the bulk of the liquid this force is exerted equally in all directions, but at the surface it makes itself felt as a contracting tendency, seeking to reduce the surface area. It is surface tension which causes a drop of liquid to take on a spherical form. It is an extremely important phenomenon in the study of emulsions and of other suspensions of minute particles. The movement of soil moisture, and of sap in plants, and the concentration of ores by the flotation process depend in large part upon surface tension.

EXPERIMENT 88—The Surface Tension of Water Supports a Safety Razor Blade.

Just as you floated a needle on water in a previous experiment, you can float a razor blade. Simply lay it gently with dry fingers on the clean surface of water in a bowl. A floating razor blade will sometimes hold up a weight of as much as a gram without sinking. You will find it interesting to touch the water at one side of the blade with a piece of soap and watch the blade be pulled away by the greater surface tension on the opposite side.

EXPERIMENT 89—A Floating Compass.



Rub a sewing needle across the pole of a magnet several times, always in the same direction. This will magnetize it. Now place it carefully on the surface of water in a bowl, and it will float. Watch it as it swings around so as to point north and south. Your compass will be a little more stable if you float it on a thin disc or cork instead of directly on the water.

EXPERIMENT 90—A Surface Tension Boat.

Make a small flat boat 2 or 3 inches long from light wood, such as cigar box lumber. Make a narrow "V" shaped slit in the stern of the boat. Insert a small piece of soap in the "V" shaped slit and carefully place your boat on the surface of clean water in a bath tub. The reduction in surface tension around the soap at the stern is so much that the tension at the forward end is able to pull the boat forward with surprising speed.

EXPERIMENT 91—Weakening the Surface Film with Soap.

Partly fill a saucer or small bowl with water, and add several measures of Sodium Carbonate (No. 4), stirring with your glass rod until the chemical is dissolved. This will soften the water so that its hardness will not destroy the soap you are going to add. If your water is very hard, it may require still more sodium carbonate.

Now float a needle or safety razor blade on the water as before, and touch a piece of soap to the surface, first on one side of the object, then on the other. Very soon the surface film will be weakened so much that the object will break through and sink.

EXPERIMENT 92—Weakening the Surface Film with Oil.

Put a drop of vegetable oil, such as cottonseed oil, on the surface of water in a saucer or bowl and see whether you can make an object float upon it. You will find that the film is now too weak to support the object, since the oil has lowered the surface tension.

This is why oil is spread on the surface of stagnant pools to prevent the breeding of mosquitoes. The larvae, wigglers, live in water, but must stick their breathing tubes through the surface to get air. Ordinarily they hang on the surface film, but when there is oil on the water, the surface tension is lowered so greatly that the surface film can no longer support them, and they sink to the bottom and die.

EXPERIMENT 93—The Surface Tension of Glass.

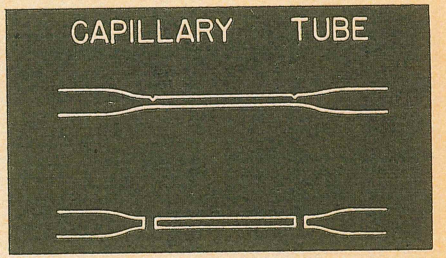
Cut a 3-inch piece from one of your glass tubes by making a scratch with a three-cornered file and then applying pressure on each side of the scratch with your thumbs and fingers. The cut end of the glass has very sharp edges. Hold it in the flame of your alcohol lamp, and when it begins to glow, you will see that the edges slowly become rounded and smooth.

Heating the glass makes it soft enough so that the surface tension can pull it into a rounded shape like a drop of liquid. Surface forces are present even in rigid solids, but they are not able to produce a change of shape until the solid is softened by heat.

Another effect of surface tension is to pull a liquid into narrow openings. If the liquid adheres to the material of which a narrow tube is made, that is, if it wets it, it will tend to creep along the surface. This increases the surface area of the liquid. Now the surface tension is always on the lookout for a chance to decrease the amount of surface, and so it tries to drag the main body of the liquid along after the creeping surface film. That is why the liquid penetrates into the tube. This is called capillary attraction.

EXPERIMENT 94— How Water Rises in a Narrow Tube.

Hold the middle of a piece of glass tube several inches long in the flame of your alcohol lamp until it softens and colors the flame yellow. Rotate the tube while you are heating it so that it will soften uniformly on all sides. Now remove it from the flame and while it is still soft pull it out several inches, constricting the central softened part to a narrow tube.



After your drawn-out tube has cooled, cut out the narrow part by scratching it near the end with a three-cornered file and breaking it. Hold the fine tube upright and dip the lower end into water in a dish. Notice that the water rises in the tube far above the level in the dish. The height of rise is greater the smaller the tube and the greater the surface tension. This is one of the ways of measuring the surface tension of liquids.

EXPERIMENT 95—Rise of Soap Solution in a Narrow Tube.

Soften half a glass of water by adding several measures of Sodium Carbonate (No. 4) or a little washing soda to it. Stir soap shavings into it until you get a good lather. Now dip into this soap solution the same end of the same tube which you used in the preceding experiment and immerse it to the same depth. Notice that the soap solution does not rise nearly so high in the tube as the water did. This is because the soap has decreased the surface tension of the water and it is therefore unable to pull itself up so far into the tube.

Many other substances besides soap have the effect of lowering the surface tension of water, and some of them are of great practical importance because the lower the surface tension of the water, the easier it is for it to wet solid surfaces. For example, in spraying leaves and fruit to kill insects and fungi, the farmer wants the liquid to stick to the leaf or fruit as a thin film, rather than to run off in drops. He accomplishes this by including in the spray mixtures a substance to lower its surface tension, thus decreasing the tendency of the liquid to gather into drops.

EXPERIMENT 96—Water Rises in a Lump of Sugar.

Put 1 or 2 teaspoonfuls of water in a saucer and add a few drops of ink. Stir this mixture until the water is uniformly colored. Now set a sugar lump on end in the colored solution, and watch how quickly the liquid rises through the sugar, being drawn up through the tiny spaces between the sugar grains just as the water was drawn up in the narrow tube used in the preceding experiment. After a time, of course, the extremely small crystals which cement together the large grains will dissolve out and the lump will crumble.

The spreading of liquids through blotting paper or a sponge is explained by surface tension in just the same way.

CONVECTION CURRENTS

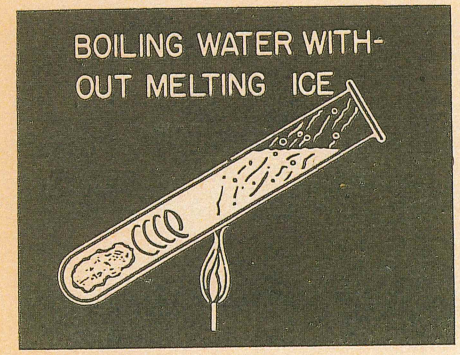
Hot water has a lower density than cold water, and so if you heat the water at the bottom of a vessel, it will rise, and the colder, denser water at the top will be forced down to take its place. You can show this very easily in a test tube.

EXPERIMENT 97—Convection Currents.

Hold the side of a test tube ¾ full of water against a flame and then add a small drop of ink. You will first notice that the color sinks to the bottom of the test tube. Then on continued heating, streamers rise up in the tube, and sink to the bottom again. A complete cycle will be described as the heated water rises, is cooled and sinks, and again rises on being reheated.

If the upper layers of water are heated instead of the lower, no convection currents can be produced, since the warm water at the top cannot rise any further.

EXPERIMENT 98—Boiling Water without Melting Ice.



Fill a test tube nearly full of ice cold water. Obtain a small piece of ice which will just fit in the test tube, and also a weight formed from a small coil of wire which you can put on top of the ice to make it sink to the bottom of the test tube. When you have your ice in the bottom of the test tube, heat the water in the upper portion of the tube until it boils. The ice will not melt for there are no convection currents to bring warm water into contact with it.

Convection currents are possible in gases as well as in liquids, and everyone is familiar with the rising current of hot air above a radiator or stove. The winds are essentially convection currents, produced by unequal heating of different areas of the earth's surface.

Convection currents play an important part in the heating of houses, for not only is the heat distributed from the radiators to all parts of the room to a large extent by convection currents, but in the hot water type of heating system, the heat is carried from the furnace to the radiators by convection currents in the water.

TESTING WATER

All water as it occurs in nature contains more or less dissolved impurities. These are usually mineral salts, but animal and vegetable substances may also be present. Water which is to be used for domestic purposes should be examined, particularly for organic matter (that is, animal or vegetable matter). It is not that this organic matter is itself necessarily harmful but it is usually accompanied by the bacteria of putrefaction and disease.

The color, odor and taste of water are not in themselves indications of its purity, but these considered with other data help in forming an opinion of the water in question. A clear, colorless, tasteless water may be polluted, while on the other hand, a safe water may have acquired color from dissolved iron or a "peaty" taste from swamp vegetation.

EXPERIMENT 99—The Odor Test.

Fill a test tube about two-thirds full of water which you desire to test. Close the mouth of the tube with your thumb and shake thoroughly for a few moments. Now remove your thumb and notice the odor of the water. Warm the water gently over a flame and again shake and notice the odor. The odor will be strengthened by heating. Always keep your face away from a test tube while heating it and remove the tube from the flame before smelling.

A putrid or offensive smell probably indicates sewage contamination.

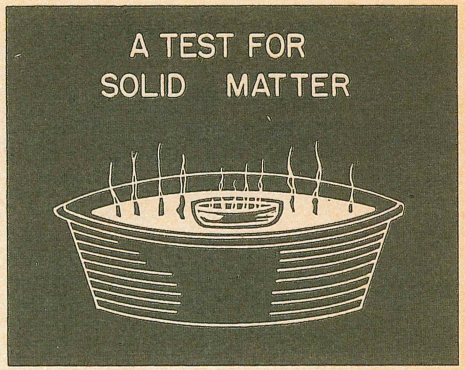
EXPERIMENT 100—The Color Test.

Fill a test tube with the water to be tested, and hold the tube in front of a piece of white paper. Observe the color and clearness of the water by this method.

If it is possible to obtain a sample of distilled water, the water to be tested should be compared with the clearness of this sample.

EXPERIMENT 101—A Test for Solid Matter.

Put a small quantity of the water to be tested in a clean china or porcelain bowl and float the bowl in a pan of boiling water. The water in the bowl will gradually evaporate, and when the bowl is dry, examine carefully for any trace of solid matter, which if contained in the water will be left as a residue in the bowl.



EXPERIMENT 102—Testing for Alkalinity.

Moisten a piece of red Litmus Paper (No. 18) with a drop of the water which you are testing. If the red Litmus paper turns blue, the water is slightly alkaline.

EXPERIMENT 103—Testing for Acidity.

Moisten a piece of blue Litmus Paper (No. 18) with a drop of the water which you are testing. If the blue litmus paper turns red or pink, the water is slightly acid.

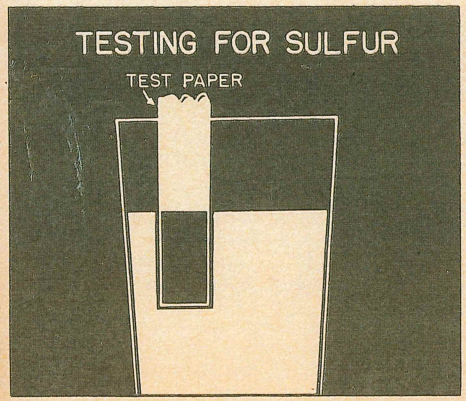
EXPERIMENT 104—Testing for Sulfates.

To a test tube of water add 1 measure of Strontium Chloride (No. 10) and shake it until dissolved. If the water, after 10 minutes, shows a white turbidity, it contains a considerable amount of sulfates, generally calcium or magnesium sulfate. Such water is usually hard.

EXPERIMENT 105—Testing for Iron.

To a test tube full of water add 1 measure of Sodium Ferrocyanide (No. 6) and shake it until dissolved. If the water shows a slight blue coloration either at once, or after standing awhile, iron is present.

EXPERIMENT 106—Testing for Sulfur.



Sulfur is present in some mineral waters in the form of hydrogen sulfide. It may be detected even in very small quantities by dropping into the water a small piece of Sulfide Test Paper (No. 31). If hydrogen sulfide or other sulfides are present, the paper will turn brown or black.

EXPERIMENT 107—Testing for Carbon Dioxide or Carbonates.

Make up a solution of lime water by adding ½ measure of Calcium Oxide (No. 20) to ½ test tube of water. Shake the tube well and let it settle for a few minutes. Add a few drops of the clear portion of this solution to a tube of the water which is to be tested. If a white turbidity or precipitate is formed, the water contains carbon dioxide, either as a dissolved gas or in the form of a carbonate or bicarbonate.

EXPERIMENT 108—Testing Water for Calcium and Magnesium.

To a test tube half full of the water to be tested add 1 measure of Sodium Carbonate (No. 4) and shake the tube to dissolve the solid. If, after standing a little while, the water shows a white turbidity or a precipitate, calcium or magnesium salts are present.

Now add 2 measures of Ammonium Chloride (No. 9) and shake the tube again. If the white precipitate which was formed by the sodium carbonate is dissolved by the ammonium chloride, the original water contained magnesium. If the precipitate does not dissolve, calcium was present, probably in the form of calcium sulfate or calcium bicarbonate. Ordinarily, hard water contains considerable quantities of calcium salts accompanied by a lesser amount of magnesium salts.

HARD WATER

The hardness of water is caused by dissolved calcium and magnesium salts, which react with soap to convert it into insoluble and useless compounds. The waste of soap caused by hard water amounts, in some communities, to many thousands of dollars a year. In steam boilers, and in heating systems, the salts in hard water tend to precipitate out as a hard scale, coating the inside surfaces of boilers and pipes. This hinders the transmission of heat and the flow of the liquid.

There are two general types of hardness which may be present in water. The first is bicarbonate hardness, caused by calcium or magnesium bicarbonates. These two compounds are converted by boiling into calcium or magnesium carbonate, substances which, being insoluble, precipitate out of the water. Therefore, bicarbonate hardness can be removed by boiling the water. For that reason it was formerly called temporary hardness.

Non-carbonate hardness is the second type, and is caused by calcium or magnesium sulfates. Since these compounds are not precipitated by boiling, water containing them was formerly called permanently hard. Both types of hard water, however, are very readily softened by certain chemicals. In the household, washing soda (sodium carbonate), ammonia water (ammonium hydroxide), and trisodium phosphate are often used. All of these precipitate the calcium and magnesium in the water by converting them into insoluble compounds.

A better softening method makes use of a mineral known as zeolite. This substance is a sodium aluminum silicate, and when hard water comes in contact with it, calcium and magnesium ions dissolved in the water are exchanged for the sodium of the mineral and are held fast on the surface of many tiny pores of the solid. Thus, when hard water containing calcium sulfate is allowed to trickle through a bed of zeolite, the calcium sulfate is converted into sodium sulfate, which does not react with soap. Hence, the water is softened.

Of course there is a limit to the amount of calcium and magnesium which the zeolite can remove from water by exchange with sodium, but when its exchange capacity is exhausted, it is a simple matter to restore its activity. This is done simply by washing it with a concentrated salt solution, which puts sodium ions back into the zeolite and removes the calcium and magnesium which the mineral has taken up. After the excess salt has been washed out, the zeolite is back in its original state and can be used to soften another quantity of water. This cycle of use and regeneration can be repeated indefinitely, and the zeolite never needs to be replaced.

A more recent development in water treatment uses various specially designed synthetic resins. These are chemically very different from the zeolite, but they have the same property of exchanging one kind of ion for another. For ordinary water softening they are made to remove calcium and magnesium ions from solution, replacing them with sodium ions. Like the zeolite water softeners, these resins can be used over and over again by simply regenerating them with a concentrated salt solution. This method of water softening is now used in many homes.

Other resins of the same general type will do even better than this and remove from water nearly all ions, both positive and negative. The result is an extremely pure water, such as formerly could be obtained only by the expensive process of distillation. Such ion exchange resins are being used more and more to supply pure water for industrial processes and chemical laboratories.

EXPERIMENT 109—The Effect of Hard Water on Soap.

Make some artificial hard water by adding ½ measure of Epsom salts (magnesium sulfate) to a test tube half full of water. Shake the tube until the solid is dissolved. Now make some soap solution by shaking 2 measures of fine soap shavings in a test tube half full of warm water. Add the soap solution drop by drop to the hard water and notice the curdy precipitate which forms. This is a magnesium soap, which is insoluble in water. It is of no use for washing.

EXPERIMENT 110—Testing Water for Hardness.

Add 2 measures of fine soap shavings to a test tube half full of water and after shaking it well, warm gently to dissolve the soap completely. Fill another test tube half full of the water to be tested and add to it 2 or 3 drops of the soap solution. Close the test tube with your thumb and shake it vigorously. If a permanent foam or suds forms, the water is soft. If not, add a few drops more of soap solution and shake the tube again. Continue this until a foam is produced. The amount of soap solution required to produce a foam is an indication of the hardness of the water. Try this with rain water collected in a clean dish.

EXPERIMENT 111—Softening Hard Water.

Prepare some artificially hard water by shaking ½ measure of Calcium Oxide (No. 20) in a test tube ⅔ full of water. Let the undissolved solid settle to the bottom of the tube and pour off a little of the clear liquid into another test tube. To this portion add several drops of soap solution prepared by shaking a few soap shavings with water. Shake the tube vigorously after you add the soap solution and you will find that no suds are formed. If you continue adding soap solution until suds form when the tube is shaken, you will realize that the water was very hard since a large amount of soap solution was required. Furthermore, you will notice a curdy precipitate of calcium soap.

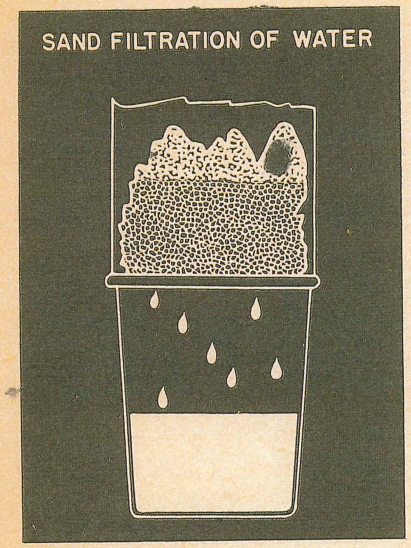
Now pour another portion of your artificially hard water into another test tube and add 1 measure of Sodium Carbonate (No. 4). Shake the tube and notice the turbidity caused by the calcium carbonate which is precipitated. The calcium hardness has now been removed from the water and when you add a few drops of soap solution you will find upon shaking the tube that a good lather or suds forms at once.

FILTERING AND CLARIFYING WATER

Many natural waters in addition to dissolved impurities contain some matter in suspension. Most cities which must use muddy water make use of a filtration process to clarify their water.

While the compounds causing hardness are dissolved and must be removed by chemical means, suspended particles which make water appear cloudy or muddy can be removed by filtration. This is done on a large scale by allowing the water to percolate through beds of sand.

EXPERIMENT 112—Filtering Water Through Sand.



Obtain a small, clean tin can and with a nail, punch about ten small holes in its bottom. Now fill it nearly full of fine clean sand. It may be well to wash the sand first by stirring it up in a pail of water and allowing it to settle. Set the can of sand on top of a clean tumbler.

Get some muddy water, or if need be, make it by stirring up a little clay or earth with water. Pour this into your sand filter, and soon clear water will begin to drip from the bottom.

Often a chemical method of clarification is applied before the water is filtered. Aluminum sulfate is added in small quantities and if the water is not sufficiently alkaline to precipitate aluminum hydroxide directly, a little lime or sodium carbonate is needed also. The hydroxide formed is a flaky, gelatinous precipitate and as it settles out, it entangles and drags down with it suspended particles and bacteria, leaving the water much clearer.

EXPERIMENT 113—Clarification of Water by a Chemical Precipitate.

To a test tube nearly full of muddy water (made so if necessary by the addition of a small amount of clay) add ½ measure of Aluminum Sulfate (No. 12) and shake the tube until the solid is dissolved. Next add ½ measure of Sodium Carbonate (No. 4) and shake it again. Let stand quietly for 15 minutes. For comparison let some of the untreated muddy water stand in another test tube along side.

SOLUTIONS

When we put sugar in tea or coffee, when we wash our hands with soap and water, when we add salt to water for cooking vegetables, we are making solutions. Practically every liquid we use is a solution. Even tap water contains not only mineral matter but also gases dissolved in it.

A substance is dissolved or "in solution" when it is so finely divided, and its particles are so completely scattered throughout the liquid that they can neither be seen nor separated from the liquid by filtering. A dissolved substance may in most cases be recovered in its original state by evaporating the liquid.

EXPERIMENT 114—Making a Solution and Recovering the Dissolved Substance.

Dissolve a spoonful of common table salt (not included in outfit) in half a glass of water. Now pour this solution into a small pan or similar cooking utensil and heat on the stove until the water is driven off. Taste what you find remaining in the pan and satisfy yourself it is the same substance that you started with.

Unless the pan is heated very gently when the salt is nearly dry you will find there will be some popping due to the expansion of steam in the crystals of salt. This phenomenon is known as decrepitation. It does no harm except to scatter some of the salt.

EXPERIMENT 115—An Insoluble Substance.

Fill a test tube half full of water and add 1 measure of Sulfur (No. 1). Close the mouth of the tube with your thumb and shake it thoroughly. The sulfur will not dissolve but remains in a solid state scattered throughout the water.

To prove that none of the sulfur has dissolved in the water, let the contents of the test tube settle and put some of the clear liquid in a clean tin cup and heat very gently until it has entirely evaporated or boiled away. There will be no solid remaining as when you evaporated the salt solution.

Most chemicals react much more readily if they are dissolved in water, as this permits the particles or molecules to come into closer contact.

EXPERIMENT 116—Separating a Soluble from an Insoluble Substance.

Place in a test tube 3 measures of Ammonium Chloride (No. 9) and 2 measures of Sulfur (No. 1) and shake the tube until they are well mixed. In order to separate them again, fill the test tube two-thirds full of water. Since the ammonium chloride is soluble in water, a little shaking will dissolve it, but the sulfur is insoluble and remains as a solid suspended in the solution of ammonium chloride.

Now filter this mixture through a piece of soft paper as described in "The Use of Chemical Apparatus." If you have worked carefully, the filtrate will be clear and free from suspended sulfur. Pour it into a saucer and set it in a warm place. When the water has entirely evaporated you will find remaining a white deposit of ammonium chloride. Remove the filter paper and spread it out to dry. On it you will find the sulfur you originally mixed with the ammonium chloride.

EXPERIMENT 117—Effect of Solution on Chemical Reaction.

Mix together on a sheet of paper 1 measure of Sodium Carbonate (No. 4) and 2 measures of Sodium Bisulfate (No. 7). The dry powder obtained shows no signs of any chemical reaction. Now put this mixture in a test tube and add a drop or two of water. Immediately a reaction occurs liberating a large quantity of gas. This gas is carbon dioxide.

EXPERIMENT 118—Another Example Showing Color.

Mix together on a sheet of paper 1 measure of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21). Here again the dry powder obtained shows no sign of any chemical reaction. Put this mixture in a test tube and fill the test tube ¼ full of water. Immediately an intense blue color is formed showing that the two substances have reacted forming still another substance.

The particles of a dissolved substance have the property of spreading themselves uniformly throughout the entire volume of a liquid. This scattering is known as diffusion.

EXPERIMENT 119—Diffusion.

To a glass ¾ full of water add 1 measure of Ferric Ammonium Sulfate (No. 21) and stir until it is completely dissolved. Now get a piece of tissue paper about 2 inches square and in its center, place 1 measure of Sodium Ferrocyanide (No. 6). Fold up the solid tightly in the paper and twist the corners together. Put the paper wad in the glass of water and set it in a quiet place. When the liquid soaks through the paper, blue streamers will descend from it, and eventually the forces of diffusion will make the liquid a uniform blue color.

Gases as well as solids tend to dissolve in liquids. It is oxygen dissolved in water which makes it possible for fish to live. Pollution by sewage and industrial wastes tends to remove the dissolved oxygen from rivers, but as more oxygen dissolves from the air, the organic waste materials are slowly oxidized, and the water is purified. A stream with many rapids and waterfalls will purify itself more quickly than one which runs quietly since oxygen will have a better chance to dissolve in it.

EXPERIMENT 120—Dissolving a Gas in Water.

Put 2 measures of Sodium Carbonate (No. 4) and 2 measures of Sodium Bisulfate (No. 7) in a test tube and fill one-quarter full of water. Quickly fit the test tube with your gas delivery tube and stopper, and insert the other end of the delivery tube in the mouth of an empty test tube so that the gas which is coming from the delivery tube will flow into the test tube just as if water were coming through the delivery tube instead of carbon dioxide gas, which you cannot see.

After a minute or so the test tube should be full of gas. Keep your finger over the mouth as much as possible and carefully fill the tube one-quarter full of water. Now press your finger or thumb tightly over the mouth and shake the tube. Do you feel a suction on your thumb showing that the gas in the tube has dissolved in the water forming a partial vacuum?

Gases always become less and less soluble the higher the temperature. Therefore, if a liquid contains the greatest possible amount of gas dissolved in it at low temperature, some of this gas must come out of the solution when the liquid is heated.

EXPERIMENT 121—The Effect of Temperature on Solubility of a Gas.

Fill a test tube half full of water drawn from the cold water faucet, and heat it slowly and gently over a flame. Notice that as the water becomes warm, bubbles of gas appear on the sides of the tube. This is air which was dissolved in the water at the lower temperature. By the time the water is nearly up to its boiling point, the sides of the tube will be thickly covered with bubbles. Boil the water for a few seconds, and then cool the tube in cold water. Again heat the water slowly and gently, and this time you will see no bubbles, or at most very few, since the previous heating removed all the dissolved air.

Now cool the test tube once more, and shake it vigorously, thereby giving some more air a chance to dissolve in it. Upon heating it again, you will find that bubbles form just as they did on the first heating.

When chemists want to remove a gas from a liquid, they usually boil the liquid.

EXPERIMENT 122—Removing a Gas from Solution by Boiling.

Place 1 drop of household ammonia in a test tube ¼ full of water. Now add 1 drop of Phenolphthalein Solution (No. 22). This will color the liquid in the test tube red, as ammonium hydroxide is a base. Heat this red liquid for a few minutes and notice that as it begins to boil the red color becomes fainter. By boiling for several minutes, the liquid may be turned colorless, but be careful not to let the hot liquid spurt out of the test tube.

Ammonium hydroxide is formed by dissolving ammonia gas in water. When the solution is boiled, the ammonia is driven out.

EXPERIMENT 123—Solubility of Liquids in Each Other.

Try to mix a few drops of light oil with a little water in a test tube. You can shake it as much as you wish, but the oil will always separate from the water. In other words, it is insoluble in water.

Now add a few drops of glycerine to a test tube ¼ full of water. As soon as you begin to shake the tube, the glycerine immediately dissolves in the water, showing that glycerine is soluble in water.

Many salts when dissolved absorb heat. Thus it is possible to lower the temperature of water by dissolving certain salts in it.

EXPERIMENT 124—Lowering Temperature by Solution.

Fill a test tube ¼ full of water and feel the temperature by holding the test tube against your cheek for a few seconds. Now add 4 measures of Sodium Thiosulfate (No. 16) and shake the tube for a moment. Bring the bottom of the tube in contact with your cheek again and notice the decided difference in the temperature.

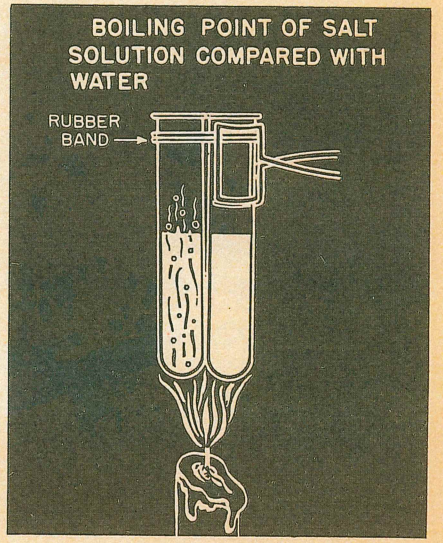
EXPERIMENT 125—Raising Temperature by Solution.

Place 5 measures of Sodium Carbonate (No. 4) in a dry test tube. Warm the tube over your alcohol lamp flame and you will notice that water of hydration is given off in the form of steam. Continue warming the tube until no more water is given off. Now set this tube aside to cool. After about 15 minutes when the tube has become cool, fill another test tube ¼ full of water and carefully notice its temperature by holding the bottom of the tube against your wrist. Now pour the water in this tube into the tube containing the dehydrated sodium carbonate. Close the mouth of the tube with your thumb and shake it. After the sodium carbonate has dissolved, again check the temperature of the bottom of the tube and notice if the water has become warmer by dissolving the dehydrated sodium carbonate.

Dissolving a salt in water raises the boiling point of the water. For example, a solution containing 25 parts of sodium chloride in 100 parts of water boils at 223° Fahrenheit (106° Centigrade). Pure water, of course, boils at 212° F. (100° C.).

EXPERIMENT 126—Boiling Point of Salt Solution Compared with Water.

Dissolve 10 measures of common salt in a test tube half full of water, and fill another test tube half full of plain water. Now heat these two test tubes over a flame. Hold them so that each will get just the same amount of heat. You will find that the water will boil before the salt solution does. Hold the tubes carefully so the boiling water will not spurt out on you.



Dissolved solids not only raise the boiling point of water, but lower its freezing point as well. This explains how salt melts ice on slippery sidewalks, and why a mixture of ice and salt is used to freeze ice cream. Saying that salt lowers the freezing point of water is the same as saying that it lowers the melting point of ice. Suppose that we add salt to a mixture of ice and water which is just at the freezing temperature. The salt coming in contact with the ice lowers its melting point below 32 degrees Fahrenheit, but since the mixture of ice and water is at just 32 degrees, the ice is warmer than its melting point and so it begins to melt. In doing so it takes up heat from the water and anything else in contact with it just as a cake of melting ice takes up heat from your hand. The ice in the mixture will continue to melt and take up

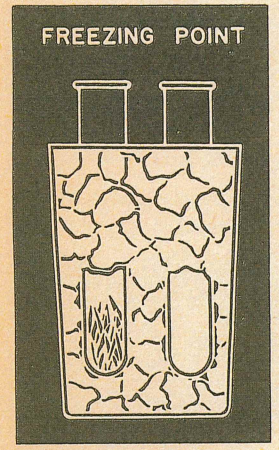
heat until it has cooled the whole mixture down to the freezing point of the salt solution. Of course if water or an ice cream mix is placed in this freezing mixture it will be frozen.

EXPERIMENT 127—The Action of Freezing Mixtures.

Crack enough ice to fill a glass or tumbler ¾ full. Now put about ¼ glass of salt in with the ice and mix the two together. Fill a test tube half full of clear water and set the tube in the mixture of salt and ice for a few minutes so that it is surrounded by the freezing mixture. Be careful that you do not allow any of the salt to get into the test tube. In a short time the water in the test tube will be frozen.

Now you can prove directly that dissolved salt lowers the freezing point of water.

EXPERIMENT 128—Freezing Point of Salt Solution Compared with Water.



Prepare a freezing mixture by mixing ice and salt just as in the preceding experiment. Fill a test tube half full of water and add 8 or 10 measures of common table salt. Shake the tube until the salt is dissolved. Fill another test tube half full of pure water. Now put both of these test tubes into the glass containing the mixture of ice and salt, and you will soon find that the pure water will freeze, while the salt solution will not.

Alcohol, as well as salt, lowers the freezing point of water, and for that reason it is used in automobile radiators in the winter. People do not put salt in their radiators to keep them from freezing because the salt increases the corrosion of metal.

Although a liquid ought to freeze when cooled down to its freezing point, it sometimes happens that it remains liquid even when the temperature is far below the freezing point. A substance which, although it is below its freezing point, is still liquid is said to be undercooled. An undercooled liquid is in an unstable condition, and sometimes the slightest jar will cause it to solidify.

EXPERIMENT 129—Undercooled Water.

Obtain half a glass of cracked ice and mix with it several spoonfuls of ordinary salt. Fill a test tube ¼ full of pure water and place it in the freezing mixture. Be careful not to allow any of the ice or salt to get into the test tube. If the water in the test tube is kept perfectly quiet, it may be cooled as much as 10° Centigrade below zero (that is, to 14° F.) without freezing. Then if it is shaken, or if a tiny crystal of ice is added, it will immediately begin to freeze and the temperature will rise to the freezing point.

We have learned that gases are always less soluble in hot water than in cold. A few solid substances obey the same rule, but most solids behave in just the opposite way, that is, they are more soluble in hot water than in cold.

EXPERIMENT 130—The Usual Effect of Temperature on Solubility.

Put 4 measures of Ammonium Chloride (No. 9) in a clean, dry test tube. Add exactly 4 drops of water. Shake the tube for several minutes and notice that the ammonium chloride will not all dissolve in this amount of water. Now warm this solution, shaking the tube as you warm it. You will see more and more of the ammonium chloride going into solution as the liquid becomes hot until finally, it should be practically all dissolved. Set the tube to one side and allow the solution to cool slowly. After awhile, crystals of ammonium chloride will form and will gradually grow larger.

EXPERIMENT 131—The Opposite Effect of Temperature on Solubility.

Fill a test tube half full of cold water and add ½ measure of Calcium Oxide (No. 20). Calcium oxide combines with water to form calcium hydroxide. Shake the tube thoroughly in order to dissolve all the calcium hydroxide possible. Now allow the tube to remain quiet for a minute, and then pour part of the clear solution into another clean test tube being very careful to leave the undissolved calcium hydroxide in the first test tube.

Heat the solution in the second tube over a flame for a few moments, and you will find that the clear solution gradually becomes slightly turbid or cloudy, since calcium hydroxide is less soluble in hot water than in cold water and part of it comes out of solution when the temperature is raised.

Sometimes it is possible to prepare a solution which contains more of a dissolved substance than is normal for the temperature. Such a solution is said to be supersaturated, and like an undercooled liquid, it is in an unstable state. Crystallization of the excess of the dissolved substance may be started by a jar or by dropping in the slightest trace of solid.

EXPERIMENT 132—A Supersaturated Solution.

Fill a test tube ⅓ full of Sodium Thiosulfate (No. 16). Add exactly 2 drops of water and warm gently. The sodium thiosulfate will melt, or dissolve in the water. Continue heating until the solid is entirely dissolved and the solution is almost boiling. Now close the mouth of the test tube with a plug of cotton and allow it to stand without being shaken until it is cold.

You now have a supersaturated solution, that is, one containing much more dissolved substance than it has any right to have at room temperature. Drop in a tiny crystal of sodium thiosulfate and at once crystals will begin to form in the supersaturated solution.

CRYSTALLIZATION

The process of crystallization is very important to the chemist, since by means of it he can purify many substances. When crystals form from a solution, they tend to reject the impurities, leaving them behind in the solution. Hence to purify a compound, we dissolve it in liquid and then allow it to crystallize out again. The crystals obtained by this recrystallization are purer than the original ones, the impurities having remained in the solution.

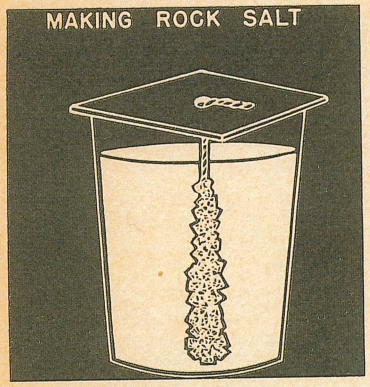
When a substance is more soluble in hot water than in cold, it is usually easy to get good crystals. All we need do is dissolve as much of the substance as possible in hot water and then let the solution cool. When cold, the solid is less soluble, so some of it must separate out as crystals. It is an invariable rule that the more slowly crystals are formed the larger they will be.

EXPERIMENT 133—The Formation of Crystals.

Put 3 measures of Sodium Ferrocyanide (No. 6) in a test tube and add 5 drops of water. Heat for a minute to dissolve the solid, then set it aside and let the solution cool slowly. After 15 or 20 minutes, a heavy deposit of yellowish crystals will be found in the test tube.

Repeat the experiment, but cool the solution rapidly as soon as the sodium ferrocyanide has all dissolved by setting the test tube in a glass of cold water. The crystals obtained this time will be much smaller than when the solution is cooled slowly.

EXPERIMENT 134—Making Rock salt.



Fill a glass tumber ¼ full of common salt and then add warm water until the glass is ¾ full. Stir the solution until the salt is all dissolved, or if necessary, add a little more water.

Cut a piece of cardboard to fit as a cover for the glass and hang a piece of string through a hole in the center of the cardboard so that it reaches nearly to the bottom of the glass. It may be necessary to tie a small weight on the string to make it sink in the solution. Now cover the glass and set it aside in a quiet place with the string hanging in the solution. Observe the glass from day to day and after a time you will notice crystals of salt forming on the string. After several days these crystals will grow quite large and you can lift them out of the glass by means of the string.

EXPERIMENT 135—Rock Candy.

Fill a glass half full of very hot water and add sugar until no more will dissolve after stirring thoroughly. Pour this clear syrup into a clean glass and fit a cover and string as in the rock salt experiment. After the sugar solution has remained standing for a few days, large crystals of rock candy will form on the string.

EXPERIMENT 136—Crystallizing A Mixture of Salts.

Place in a test tube 1 measure of Sodium Bisulfate (No. 7), 1 measure of Ferric Ammonium Sulfate (No. 21), 1 measure of magnesium sulfate (Epsom salts) and 6 or 7 drops of Cobalt Chloride Solution (No. 27). Carefully add water until the tube is about ¼ full and then heat gently until the solids are dissolved. Now pour the clear, hot solution into a clean glass, being careful to leave any undissolved solids in your test tube. Set the glass in a quiet place and cover with a small piece of cardboard. Notice after a day or two the crystals which have formed in the solution. Chemists often separate different compounds from each other by crystallization since some substances crystallize more easily than others.

EXPERIMENT 137—Frosting Glass with Crystals.

Put about ½ inch of water in a test tube and add 5 measures of Ammonium Chloride (No. 9) and only 1 small drop of liquid glue or mucilage. Now heat the solution until the solids are all dissolved and apply the hot liquid, using a small brush, to the glass which you wish to frost. The liquid will begin to evaporate almost immediately and crystals will be formed which give the glass the appearance of a frosted pane.

On a larger scale, frosting glass can be done by dissolving magnesium sulfate (Epsom salts) in hot water until no more will go into the solution. Add a few drops of liquid glue or mucilage and apply this solution to the glass to be frosted.

Water is contained in many solid dry substances in a state of chemical combination. For example, some crystals contain what is called "water of hydration" or "water of crystallization".

EXPERIMENT 138—Water of Hydration.

Drop 1 measure of Sodium Ferrocyanide (No. 6) into a clean dry test tube and heat the bottom of the tube with a candle flame. Do not hold the tube in the flame, as this will cause soot to deposit. Rather hold it about ½ inch above the top of the flame. Keep your face away from the test tube while you are heating it. Notice that the powder becomes wet and water is driven off in the form of steam.

EXPERIMENT 139—Water of Hydration in Washing Soda.

Washing soda is a good example of a substance containing water of hydration. If you are familiar with this material, you know that it sometimes comes in clear, nearly transparent, crystals while at other times it is covered with white powder, and is very dusty. This is because washing soda loses part of its water of hydration when it is exposed to the air, and the white dusty powder has probably been stored so long that much of the water has been lost. It is just as well to remember, however, that a pound of washing soda that has lost part of its water contains much more soda than a pound of the big, good-looking crystals which still have the water in them. Heat a few measures of washing soda in a test tube and notice the large amount of water which is given off.

EXPERIMENT 140—Water of Hydration in Cobalt Chloride.

Put 5 or 6 drops of Cobalt Chloride Solution (No. 27) in a clean test tube and warm this solution gently. Suddenly you will notice that the pink solution has turned blue. Stop heating and allow the test tube to cool and the solution will again be pink. Heat the tube again until the moisture is all driven off from the cobalt chloride and you will find a blue solid remaining. Now add just 1 drop of water and the blue solid again becomes pink. These changes are due to water of hydration. When heated, the water of hydration is driven off from the cobalt chloride and it becomes blue in color. When allowed to cool, the cobalt chloride again takes up moisture as water of hydration and becomes pink.

EXPERIMENT 141—A Cobalt Chloride Barometer.

Cobalt chloride is often used as a form of chemical barometer. A sheet of white paper is moistened with a concentrated solution and allowed to dry. It is then hung where the outside air will circulate freely about it, but where the rain will not reach it. When it looks bluish it is an indication of dryness and usually an indication of continued clear weather. When it is pink it indicates wet or moist weather.

Such a barometer can easily be made by painting Cobalt Chloride Solution (No. 27) on a sheet of paper with a soft brush. These barometers or weather indicators are often seen in the form of dolls or figures whose dresses are painted with cobalt chloride solution, or they can be made up in the form of pictures which change color according to the weather.

EXPERIMENT 142—Testing for Traces of Moisture.

A convenient method of testing for small traces of moisture is as follows: Prepare a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. Saturate some small strips of soft, white paper with this solution and dry the strips. When you are ready to use them, place half a measure of Sodium Ferrocyanide (No. 6) on one of the dry strips and spread it over the surface by means of a soft brush or wad of cotton.

The slightest trace of moisture produces dark blue spots on these strips.

HYDROGEN

Hydrogen is a colorless, odorless gas. As we have learned, it is contained in water in combination with oxygen. Practically all organic substances, such as wood, petroleum, sugar, fats and proteins contain hy-

drogen combined with carbon, and often with other elements. Hydrogen is an essential constituent of all the acids, and is often referred to as "the acid former".

Because it is the lightest gas known, hydrogen was sometimes used in balloons and zeppelins, but since its inflammability has led to numerous serious accidents, an example of which is the Von Hindenburg disaster, the more expensive helium has replaced it for this purpose.

Hydrogen is a rather active element, combining, for example, directly with oxygen, that is, burning, to form water. As we have already learned, ammonia is formed when hydrogen reacts with nitrogen under pressure. In the so-called hydrogenation of oils, hydrogen under pressure is caused to combine with hot vegetable oils to form solid cooking fats. A nickel catalyst assists the reaction.

One of the most common ways of preparing hydrogen in the laboratory is by the reaction between a metal and an acid. Sodium bisulfate, $NaHSO_4$, can be considered as partly sodium sulfate and partly sulfuric acid. When dissolved in water, this chemical forms the equivalent of a dilute solution of sulfuric acid. In the following experiments we will, therefore, use sodium bisulfate dissolved in water and will consider the reactions between the sulfuric acid contained in this solution and the metal.

EXPERIMENT 143—Making Hydrogen from Sulfuric Acid with Iron.

Prepare a solution of 4 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Warm the solution slightly, if necessary, to dissolve the solid. Now add 1 measure of powdered iron metal or small pieces of iron obtained by cutting fine steel wool with scissors. There will be a reaction and hydrogen gas is given off. In this experiment, the hydrogen in the sulfuric acid was replaced by the iron, forming ferrous (iron) sulfate and hydrogen.

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(iron)}}{Fe} \rightarrow \underset{\text{(iron sulfate)}}{FeSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 144—Making Hydrogen with Aluminum.

Make a solution of sulfuric acid by dissolving Sodium Bisulfate (No. 7) in water as in the preceding experiment and add a small piece of aluminum metal. You can probably obtain aluminum from an old pan or kettle which is no longer good for cooking purposes. Notice that the acid reacts with the aluminum and that bubbles of hydrogen gas form on the metal.

In this experiment the hydrogen of the sulfuric acid was replaced by the aluminum, forming aluminum sulfate and hydrogen.

$$\underset{\text{(sulfuric acid)}}{3\ H_2SO_4} + \underset{\text{(aluminum)}}{2\ Al} \rightarrow \underset{\text{(aluminum sulfate)}}{Al_2(SO_4)_3} + \underset{\text{(hydrogen)}}{3\ H_2\uparrow}$$

EXPERIMENT 145—Making Hydrogen with Zinc.

Obtain a piece of zinc from an old flashlight battery and after cleaning it so that it is bright and shiny, cut off a few small pieces. Now prepare a solution of sulfuric acid as in the preceding experiment and to the sulfuric acid solution add a few of the small pieces of zinc. You will again find bubbles of hydrogen gas given off as the hydrogen of the sulfuric acid is replaced by the zinc, forming zinc sulfate and hydrogen. In this experiment the action should be more vigorous than in the case of either the iron or the aluminum.

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(zinc)}}{Zn} \rightarrow \underset{\text{(zinc sulfate)}}{ZnSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 146—Filling Bubble Balloons with Hydrogen.

Make some soap suds by stirring a small piece of soap in a saucer of soft water. Now prepare hydrogen as you did in the preceding experiment. Be sure to use plenty of zinc and cut it into very thin slivers or pieces. Fit your gas delivery tube to the mouth of the test tube as soon as the gas begins to come off from the zinc.

Insert the end of the gas delivery tube under the surface of the soap suds in the saucer and as the hydrogen gas emerges from the tube, bubbles will be formed that will rise through the air. Each of these hydrogen-filled bubbles is in effect a tiny ballon buoyed up by the hydrogen which it contains.

EXPERIMENT 147—Displacement of Hydrogen from an Alkali by Aluminum.

Place 3 measures of Sodium Carbonate (No. 4) and 1 measure of Calcium Oxide (No. 20) in a test tube half full of water. Shake the test tube thoroughly to dissolve all of the solids that you can. Now allow the tube to stand quietly for a few minutes and then pour off the clear liquid into another test tube. Add some small pieces of aluminum to the solution in this test tube and warm slightly over a flame. You will soon notice bubbles of hydrogen gas coming off rapidly. While the hydrogen gas is being given off, bring a flame near the mouth of the tube and you will probably notice a slight puff as the hydrogen gas burns. Be sure to keep your face at a distance when you try this.

You formed a dilute solution of sodium hydroxide when you mixed the sodium carbonate and the calcium oxide. The aluminum metal displaced part of the hydrogen from the sodium hydroxide, forming aluminum hydroxide, which is called sodium aluminate.

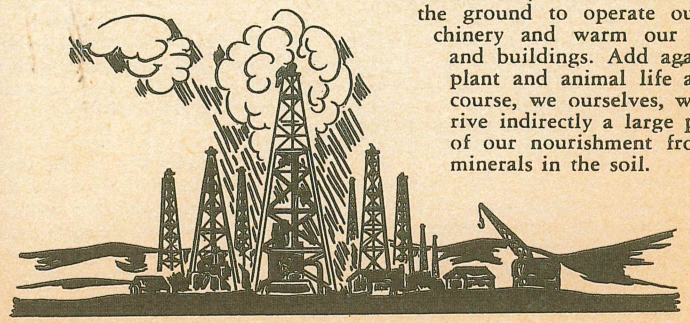
$$\underset{\text{(aluminum)}}{2\ Al} + \underset{\text{(sodium hydroxide)}}{6\ NaOH} \rightarrow \underset{\text{(hydrogen)}}{3\ H_2\uparrow} + \underset{\text{(sodium aluminate)}}{2\ Na_3AlO_3}$$

The Chemical Wealth Within Our Earth

We have learned about the atmosphere which surrounds our earth and about the water which is on our earth, so we will now take up the earth itself. What is in our earth? The surface may consist of rocks or a very thin layer of clay or sand or gravel, usually covered with a few inches of loam, or top soil. Underneath we have a variety of materials, mostly classified as minerals and rocks.

A mineral is a natural inorganic substance, usually having a definite chemical composition and often a characteristic crystalline form. Rocks are large masses of minerals or mixtures of minerals. They do not have a definite chemical composition.

Minerals are our world's greatest source of wealth. Consider the value of the iron, copper, lead, zinc, aluminum; and the precious metals, gold, silver and platinum; the diamonds and other gems. Add the value of the coal and oil, which we take from the ground to operate our machinery and warm our homes and buildings. Add again the plant and animal life and, of course, we ourselves, who derive indirectly a large portion of our nourishment from the minerals in the soil.



From a chemical standpoint, minerals may be either impure metallic or non-metallic elements, or compounds of elements. In either case, we must purify and convert them into the desired form. In the following pages we will take up many of the more common elements and learn from what minerals they are obtained and by what processes they are converted into usable form.

THE PERIODIC SYSTEM

Chemists have classified all of the elements into various groups and periods according to their atomic numbers. This classification, known as the Periodic Table, reveals certain striking similarity and relationships between the elements in the same periods and in the same groups. Thus knowing, say, the physical and chemical characteristics of chlorine and then referring to the Periodic Table, we can predict accurately the physical and chemical characteristics of the other elements in that same group, namely fluorine, bromine and iodine. It is found that the elements in a certain group exhibit similar valence characteristics, as well as similar arrangement of electrons in their atomic structure. The periodic system has become an important tool of the research chemist.

Thus far we have taken up the elements oxygen, nitrogen and hydrogen, all of which are gases under ordinary conditions of temperature and pressure. Hydrogen, we find, is in Group Ia of the Periodic Table and has a valence of one. Nitrogen is in Group Vb of the Periodic Table and may exhibit any one of five different valences, namely 1, 2, 3, 4 or 5. Oxygen, we find, is in Group VIb of the Periodic Table and this element exhibits a valence of 2. It is too soon to try and draw any comparisons at this point, but as we take up other elements, we will soon begin to see the relationships which are brought out by the Periodic Table. You will find it interesting to refer to your chart of the Periodic Table as you study the various elements.

THE HALOGENS

The elements fluorine, chlorine, bromine and iodine are so similar in their properties and reactions that they form a natural family called the halogen group of elements. These non-metallic elements occur in Group VIIb of the Periodic Table and all of them exhibit a valence of one.

Fluorine is a pale yellow gas, exceedingly active and very poisonous. Chlorine is a greenish yellow gas, very soluble in water, heavier than air and also poisonous. Bromine is a dark reddish liquid, quite destructive to tissue. Iodine is a dark gray, lustrous solid. It is almost insoluble in water, but it is quite soluble in alcohol.

The halogens are so active chemically that they are never found free in nature. They are always combined with some other element or elements. The word "halogen" means "salt former". In combination with other elements, the halogens are among the most important compounds, both industrially and in our everyday life. For example, our common table salt, sodium chloride, is a combination of sodium and chlorine.

FLUORINE

Fluorine is the most reactive of the halogens and indeed it is the most active element known. Few substances are so stable that they are not attacked and decomposed by fluorine. The minerals fluorspar and cryolite are the chief sources of fluorine compounds. The element fluorine itself is very difficult to prepare as no chemical reagent is sufficiently powerful to liberate it from its compounds. To prepare pure fluorine, we must resort to the electric current.

Fluorspar (calcium fluoride) is the source of the interesting compound hydrofluoric acid, H_2F_2, which dissolves glass. It is used for frosting and etching glass and of course, must be kept in rubber, wax or lead bottles.

Cryolite is important in the aluminum industry. It is a complex sodium and aluminum fluoride.

The difficulties and dangers involved in preparing and handling fluorine greatly retarded the development of commercial applications. Recently, however, improved techniques have made it available on a large scale and it is now being used in the manufacture of various compounds which can not be obtained in other ways. The field of usefulness of fluorine will probably be greatly widened within the next few years.

CHLORINE

Chlorine is by far the most abundant and the most important of the halogens. It occurs widely distributed in the form of various chlorides, the most common of which is sodium chloride. Since most chlorides are readily soluble in water, it is not surprising that they have been washed into the ocean in large amounts and that our chief source of chlorides is the salt beds left by the evaporation in past ages of land-locked seas.

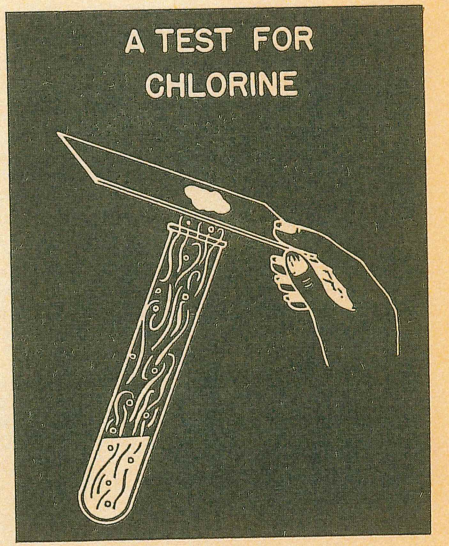
Chlorine gas is usually made from sodium chloride by an electrochemical process. The gas is liquefied by pressure and shipped in steel cylinders and tank cars for use in bleaching paper and textile fibers, in purifying water, and in many branches of chemical industry.

EXPERIMENT 148—How to Make Chlorine.

Place 1 measure of Bleaching Powder (No. 38) in a test tube, add about ½ inch of water, and shake vigorously, closing the mouth of the tube with your thumb. Now add 3 measures of Sodium Bisulfate (No. 7) to the white suspension in the tube. At once bubbles of gas begin to come off, and after a moment you can smell the characteristic odor of chlorine at the mouth of the tube. Be careful not to smell much of it, since it irritates the nasal passages.

EXPERIMENT 149—A Test for Chlorine.

Moisten the dull side of an exposed photographic negative and expose it to chlorine gas by holding it at the mouth of a test tube where chlorine is being generated. The minute particles of silver metal which form the dark image on the negative react with chlorine forming silver chloride. This is a white substance, and hence the dark portions of the negative take on a milky or cloudy appearance.



Bleaching powder is prepared by allowing chlorine gas to react with lime, calcium hydroxide, in large chambers. When mixed with water it forms calcium hypochlorite along with other compounds, and as you have just seen, chlorine is evolved when an acid is added to this mixture. Chlorine reacts slowly with water, liberating oxygen. At the moment when the oxygen atoms are first liberated, they are extremely active chemically, but after a small fraction of a second, they have all combined in pairs, forming oxygen molecules, O_2, in which state they pass off as gas. If we put

colored materials into the mixture when oxygen is being set free, a few of the oxygen atoms will come in contact with the colored substance before they have time to pair up into molecules, and these very active oxygen atoms will react with, or oxidize, many colored materials, destroying the color. This is why bleaching powder is used for bleaching. Since it exerts a similarly destructive effect upon bacteria, bleaching powder is also a good disinfectant.

EXPERIMENT 150—Bleaching with Bleaching Powder.

Place 6 measures of Sodium Bisulfate (No. 7) and 4 measures of Bleaching Powder (No. 38) in a clean glass and add a few drops of water. Now obtain various scraps of paper, cloth, flowers, leaves, and a piece of litmus paper. Moisten these slightly and place them in the glass with the bleaching powder and sodium bisulfate. Cover the top of the glass with a saucer and let it stand for about an hour.

Some of the colors will be completely removed, others will be paler or changed to a different color, while some may not be affected at all. Printing ink and lead pencil marks owe their color to carbon and the bleaching powder has no effect upon them.

EXPERIMENT 151—Making Sodium Hypochlorite Bleach.

Fill a china cup about ¾ full of water and dissolve in it all of the sodium chloride (common table salt) that will go into solution after thoroughly stirring but without heating the solution. Take the small carbon rods from 2 old flashlight batteries and connect them, by means of wires wrapped around the rods or soldered to the brass caps, to 3 dry cells connected in series. Dip these carbons into the salt solution to serve as electrodes. Allow the current to pass for 20 minutes, stirring the solution from time to time so that the chlorine gas which is formed at the positive electrode during the electrolysis will react with the sodium hydroxide, which is produced around the negative electrode. This forms sodium hypochlorite. After about 20 minutes of operation, test the solution for bleaching properties by dipping into it a strip of Litmus Paper (No. 18).

EXPERIMENT 152—Another Method of Making Sodium Hypochlorite.

Fill a test tube about ⅓ full of water and add 4 measures of Bleaching Powder (No. 38). Shake the mixture thoroughly. In another test tube ⅓ full of water dissolve 4 measures of Sodium Carbonate (No. 4). Now pour the sodium carbonate solution into the test tube containing the bleaching powder and water. Shake this mixture for a moment and then filter. In this experiment you precipitated calcium carbonate and you now have in solution sodium hydochlorite and sodium chloride.

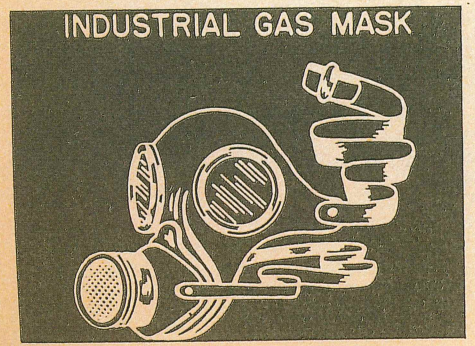
EXPERIMENT 153—Reaction of Chlorine with Sodium Thiosulfate.

Dissolve 2 measures of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. In another test tube place 2 measures of Bleaching Powder (No. 38), add ½ inch of water, and shake the tube vigorously. Then add 3 measures of Sodium Bisulfate (No. 7) and quickly fit the tube with the gas delivery tube and stopper. Pass the chlorine gas which is being evolved through the delivery tube into the sodium thiosulfate solution for a minute or two.

Sodium thiosulfate converts chlorine into sodium chloride, and while doing so it is itself changed into other substances. Among other things, free sulfur is formed, and that is the white precipitate which you see.

EXPERIMENT 154—A Gas Mask to Protect Against Chlorine.

Since sodium thiosulfate converts chlorine into a harmless compound it is an excellent protection against this irritating and poisonous gas. Chlorine was the first poison gas used in World War I and the gas masks used to protect against it contained sodium thiosulfate. Later, more deadly gases were introduced, and other chemicals had to be found to combat them.



Put 2 measures of Sodium Thiosulfate (No. 16) and 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water and shake it to dissolve the solids. Obtain a piece of cotton or linen cloth about 3 inches square and soak it in this solution for several minutes. When the cloth is thoroughly wet take it out and dry it carefully.

Now prepare chlorine from bleaching powder as in a previous experiment, and put the dry cloth over the mouth of the test tube while the chlorine is being generated. Smell at the mouth of the test tube next to the cloth and you will find that you cannot smell the chlorine gas. Put the cloth over your nose and try to smell the gas. You will find that the chemicals in the cloth react with the chlorine gas, changing it into a harmless compound.

Hydrochloric acid is one of the most important compounds of chlorine. A molecule of it is made up of one hydrogen atom and one chlorine atom, so its formula is HCl. It may also be called hydrogen chloride. Pure HCl is a gas, but it dissolves easily in water, and we usually use it in the form of a solution, called simply "hydrochloric acid". Such a solution has many uses — for cleaning steel preparatory to galvanizing or tinning, for making chlorides and numerous organic chemicals, in sugar refining, tanning, making glue, in dyeing, and in bleaching.

Hydrochloric acid is made commercially by heating sulfuric acid with common salt:

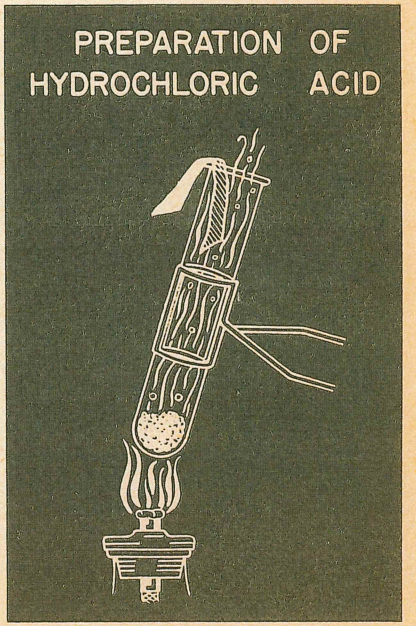
$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(sodium chloride)}}{2\ NaCl} \rightarrow \underset{\text{(sodium sulfate)}}{Na_2SO_4} + \underset{\text{(hydrochloric acid gas)}}{2\ HCl}$$

The gas which comes from the hot mixture is dissolved in water and ordinary hydrochloric acid solution is obtained. Very pure hydrochloric acid is often made by burning together hydrogen and chlorine gases:

$$H_2 + Cl_2 \rightarrow 2\ HCl$$

EXPERIMENT 155 —Preparation of Hydrochloric Acid.

In a dry test tube place 3 measures of Sodium Bisulfate (No. 7) and 1 measure of sodium chloride (common table salt). Heat this mixture strongly over a flame for a few moments, keeping your face away from the mouth of the tube. When fumes appear at the mouth of the tube, introduce a piece of moistened blue Litmus Paper (No. 18) and the hydrochloric acid gas will turn it red. Remove the test tube from the flame and smell cautiously at its mouth. You will notice the irritating odor of hydrochloric acid. Dissolving this gas in water would form the ordinary solution which we call hydrochloric acid.



Dip your glass rod into a bottle of household ammonia, and hold it near the mouth of the test tube in which hydrochloric acid is being generated. The white fumes are ammonium chloride, formed by the direct union of ammonia gas and hydrochloric acid gas:

$$\underset{\text{(ammonia)}}{NH_3} + \underset{\text{(hydrochloric acid)}}{HCl} \rightarrow \underset{\text{(ammonium chloride)}}{NH_4Cl}$$

BROMINE

Bromine, the next element of the halogen family, is a dark red liquid with a suffocating odor. When heated to about 60° Centigrade, it boils forming a choking red gas with an odor somewhat like that of chlorine. The vapors attack the eyes, nose, and throat, and the liquid can cause painful burns on the skin. Just as chlorine is less active than fluorine, so bromine is less active than chlorine, but it can form many useful compounds, both organic and inorganic.

Much bromine is obtained as a by-product from salt works. A plant on the coast of Texas extracts it directly from the sea. A ton of sea water contains only about two ounces of bromine (in the form of bromides), and 100,000 gallons of sea water must be pumped through the plant every minute to obtain a daily production of 45,000 pounds of bromine.

Bromine compounds are used in connection with anti-knock compounds in gasoline, in sensitizing photographic films, in tear-gas bombs, and in medicines.

IODINE

The last of the halogens is iodine, a grayish, lustrous solid, which is converted into a dark liquid and then into a heavy violet vapor when heated. Some iodine is obtained as a by-product in the refining of sodium nitrate (Chile saltpeter) in Chile, and also from the ashes of kelp, a kind of seaweed. Recently most of our iodine has been obtained by treatment of oil well brines, which in some localities contain from 30 to 70 parts per million of iodine in the form of sodium iodide.

Iodine is the least active of the halogens, and combines less readily with other elements. For this reason the other halogens are able to displace iodine from many of its compounds.

Small amounts of iodine in the diet are essential to human health. The thyroid gland in the neck uses it in making the complicated organic compound, thyroxin, $C_{15}H_{11}ONI_4$, which circulating in the blood, controls the rate of heat production in the tissues.

Iodine is an excellent antiseptic, being used chiefly in the form of a solution in alcohol, tincture of iodine. The compound, iodoform, CHI_3, is also a useful antiseptic. Large quantities of iodine in the form of silver iodide are used in the photographic industries, and iodine finds uses also in the dyestuffs industry and in chemical analysis.

EXPERIMENT 156—The Starch Test for Iodine.

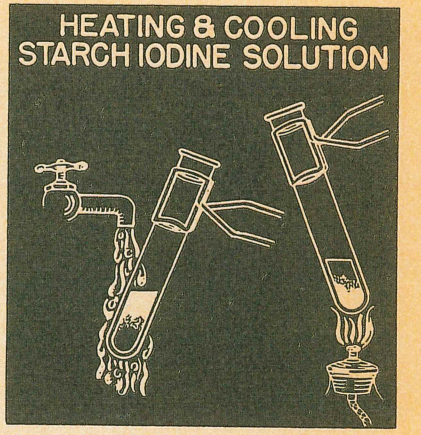
Cut a thin slice from a raw Irish potato. Secure some tincture of iodine and allow 2 or 3 drops to fall on the slice of potato. Notice the beautiful blue coloration produced.

Starch (when cold) always gives this characteristic blue coloration in the presence of iodine.

EXPERIMENT 157—The Behavior of Starch Iodine Solution.

Crush a small piece of potato (¼ inch square), place it in a test tube and add ½ test tube of water. To this add a few drops of tincture of iodine as in the preceding experiment. Add enough to produce a beautiful blue color.

Warm the test tube gently over a flame, and watch the color gradually fade. Before the liquid begins to boil, it will lose all of its color. Hold the test tube in cold water, and the blue color will slowly reappear. This experiment shows that solutions, to show the starch iodine color must be cold, for a hot solution will give no blue color.



Pour half of your blue solution into another test tube, and to one of the portions add ½ measure of Bleaching Powder (No. 38). Notice how rapidly the blue color is destroyed by the chlorine liberated from the bleaching powder.

To the other portion of the blue solution, add 4 measures of Sodium Carbonate (No. 4) and shake the tube. When a blue starch iodine solution is made alkaline, the free iodine is converted into sodium iodide and sodium hypoiodite and the blue color disappears. By neutralizing this solution with an acid, the blue color can be made to return because free iodine is again formed.

In testing for iodine or starch by means of the blue starch iodine solution, it is important to remember that you will not obtain the blue coloration: (1) if the solution is warm, (2) if the solution contains free chlorine, (3) if the solution is alkaline.

In some parts of the country the soil contains so little iodine that drinking water and vegetables do not supply enough iodine to the inhabitants, who therefore may suffer from goiter. For this reason the manufacturers of table salt add a very small amount of potassium iodide to the salt and call the product "iodized salt". Although there are only about 3 or 4 ounces of iodine in a ton of the salt, the starch test is easily able to detect it.

EXPERIMENT 158—Showing that Iodized Salt Contains Iodine.

Add about 2 teaspoonfuls of iodized salt and 5 measures of Sodium Bisulfate (No. 7) to a glass half full of water and stir until they are dissolved. Now prepare a starch solution by heating 1 measure of household starch in a test tube half full of water. As soon as the solution has started to boil, stop heating and allow it to cool. Add some of this starch solution to your salt solution. Then add a few tiny particles of Bleaching Powder (No. 38) and stir the liquid thoroughly. If a blue color, showing the presence of free iodine, does not appear, add another tiny portion of bleaching powder and stir vigorously. Continue adding bleaching powder until the liquid turns blue, but be very careful not to add too much at a time, because as you recall from a previous experiment, any excess of bleaching powder will destroy the blue color of the starch-iodine mixture. If you should add too much bleaching powder, the addition of another spoonful of iodized salt, along with, perhaps, a little more sodium bisulfate, will probably restore the blue color.

In one of our experiments with chlorine, we found that chlorine gas reacted with sodium thiosulfate. Since iodine is very similar to chlorine in chemical behavior, it is not surprising that it too reacts with sodium thiosulfate.

EXPERIMENT 159—Reaction of Iodine with Sodium Thiosulfate.

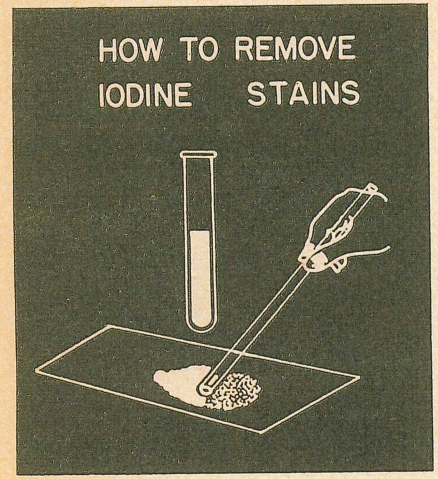
To a portion of the blue starch-iodine solution obtained in the preceding experiment, add 1 measure of Sodium Thiosulfate (No. 16) and shake it for a moment. The brown color of the water solution fades at once.

Instead of using the iodine solution left from the last experiment, you can make some more by adding 2 drops of tincture of iodine to a test tube half full of water. When the mixture has taken on a good brown color, drop in 1 measure of Sodium Thiosulfate (No. 16) and the color will disappear immediately.

The brown solution of iodine reacting with the thiosulfate forms two colorless compounds:

$$\underset{\text{(iodine)}}{I_2} + \underset{\text{(sodium thiosulfate)}}{2\ Na_2S_2O_3} \rightarrow \underset{\text{(sodium iodide)}}{2\ NaI} + \underset{\text{(sodium tetrathionate)}}{Na_2S_4O_6}$$

EXPERIMENT 160—How to Remove Iodine Stains.



This reaction is a good one to use for removing iodine stains from the skin or from fabric.

Make a solution of 4 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water, and moisten the iodine stain with this solution. If the stain is not entirely bleached after standing a few minutes, repeat the application of thiosulfate solution.

Iodine, like chlorine, can enter into combination in the molecules of organic compounds. For example, just as we have methyl chloride, CH_3Cl, we have methyl iodide, CH_3I, and corresponding to chloroform (trichloromethane), $CHCl_3$, we have iodoform (triiodomethane), CHI_3. This substance is used as an antiseptic. It is very simply made by warming alcohol with iodine in the presence of an alkali.

EXPERIMENT 161—Making Iodoform.

Add 2 measures of Sodium Carbonate (No. 4) to a test tube half full of water. Shake the tube to dissolve it and then heat the solution to boiling. Now add to the hot solution about 15 drops of tincture of iodine, which of course is simply a solution of iodine in alcohol. Almost at once the brown color of the iodine disappears, and the liquid becomes cloudy due to the fine particles of iodoform which have formed. Let the test tube stand for a few minutes and the particles will grow larger and settle to the bottom. Iodoform is a yellow solid with a peculiar odor which may remind you of a doctor's office. The iodoform can be separated from the liquid by filtering.

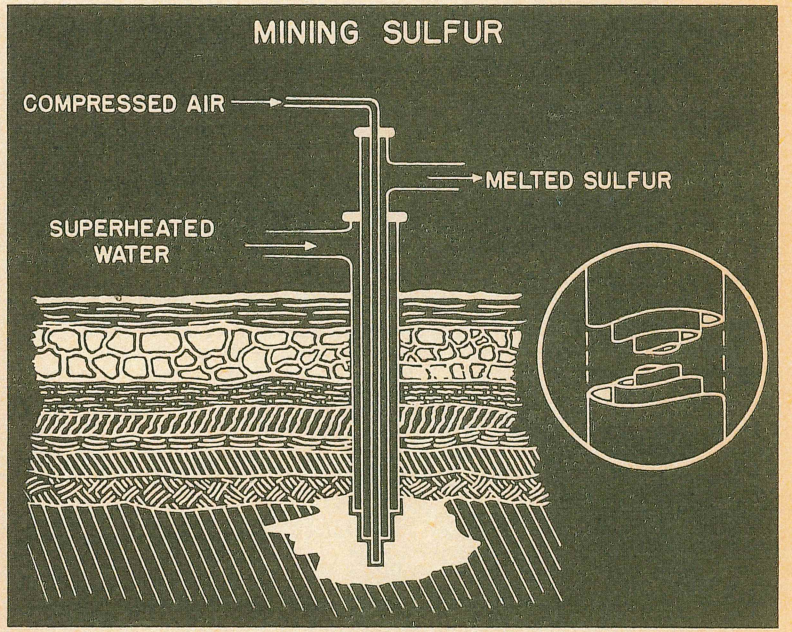
SULFUR

Sulfur is one of the world's most useful elements. You rarely see it in everyday life, but it enters directly or indirectly, into the manufacture of almost everything you use. Rubber, paper, metal, concrete, pigments, matches, medicines, and glass have all come into contact with sulfur or sulfur compounds at some stage of their manufacture. As you read the following pages, you will find many other examples of the part played by sulfur in the industries of the world.

Almost all of our sulfur comes from Texas and Louisiana, where it occurs in great deposits deep in the ground. For many years these deposits defied the efforts of mining engineers to work them in commercial quantities. Many attempts were made to obtain sulfur by digging deep shafts applying the methods used in coal mining. All of these attempts were hopeless failures because of quicksands, water, and poisonous gases.

It was not until 1903 that a successful method of tapping this great store of sulfur was put into operation. The credit for this achievement goes to Herman Frasch, an engineer who had long experience drilling wells in oil fields. He came to the conclusion after studying the problem that ordinary mining methods were out of the question, but that if the sulfur could be melted underground and pumped to the surface in a liquid condition, the problem would be solved.

As finally worked out, the Frasch process uses three pipes, one inside the other, driven down five hundred feet to the beds of sulfur. Through the outside pipe, about 8 inches in diameter, hot water is pumped down to the sulfur. This water is superheated, that is, it is heated under pressure to a temperature well above its boiling point. Coming into contact with the sulfur, this hot water melts it and forms a pool of molten sulfur at the bottom of the pipe. Hot compressed air is now forced down through the smallest pipe and, as it blows through the melted sulfur, a froth or foam of sulfur and hot water is formed. This is forced by the pressure of the hot water and the compressed air to flow up through the middle pipe to the surface. The melted sulfur is piped to large bins, where it



is allowed to solidify. About 2,000,000 tons of sulfur are brought to the surface every year by this method. It is estimated that there are more than 60,000,000 tons of sulfur under the Gulf Coast of Louisiana and Texas.

Sulfur exhibits chemical properties quite similar to oxygen, which is what we would expect, since sulfur follows oxygen in Group VIb of the Periodic Table. Sulfur is a non-metallic element, and it commonly exhibits a valence of 2, 4, or 6.

Sulfur is a very versatile substance, assuming a great variety of forms. If we carefully heat some sulfur in a test tube, it melts easily, giving a pale yellow, oily liquid. Upon heating more strongly, the liquid becomes darker and thicker, and finally it is almost black and so thick that the test tube can be turned upside down without spilling it. Heated to a still higher temperature, it becomes again more fluid and at a high temperature can be made to boil.

EXPERIMENT 162—Liquid Sulfur.

Put 3 measures of Sulfur (No. 1) into a dry test tube and heat gently over a flame, keeping your face away from the mouth of the tube. Notice that the sulfur melts at once giving a straw-colored liquid. When you heat it more strongly, the color becomes darker, and the liquid thickens so much that it will not run out of the tube. Further heating causes it to boil, and the vapor condenses on the upper part of the test tube.

Just as the element carbon occurs in three forms, diamond, graphite and amorphous carbon, so the element sulfur may exist in any one of several forms, or allotropic modifications. Ordinary sulfur consists of crystals (sometimes very small) called *rhombic sulfur*. At a temperature above 96° C. these change, more or less slowly, into another kind of crystals, monoclinic sulfur, which when cooled changes back to the rhombic form.

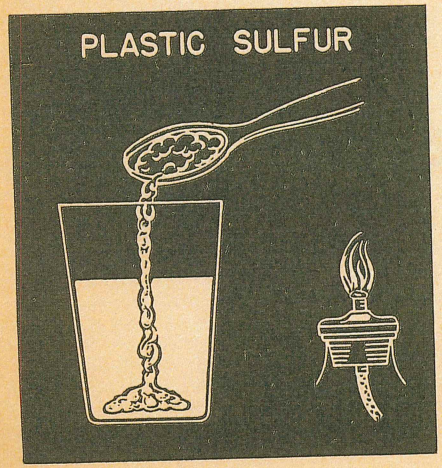
EXPERIMENT 163—Crystalline Forms of Sulfur.

Place 5 measures of Sulfur (No. 1) in your spoon and heat it gently over a candle flame until it is completely melted. Be very careful not to let it catch fire. Whenever fumes begin to come from it, remove it from the flame for a few moments. If it should start to burn, extinguish it by covering the spoon for a few seconds with a piece of metal or glass.

When the sulfur is all melted, remove the spoon from the flame and let it cool slowly. Soon the liquid will begin to solidify, or freeze. Watch closely, and when solidification is well started, tip the spoon so that the sulfur which is still liquid will run to one side, exposing the crystals already formed. Notice that these are shiny and generally somewhat elongated. They are monoclinic sulfur crystals.

Monoclinic sulfur is unstable at ordinary temperatures and tends to change over to the rhombic form, though the change may occur very slowly. Observe your monoclinic sulfur at intervals of half an hour, and you will soon see patches of dull yellow, which, after several hours, will spread over the whole of the material. Each of the monoclinic crystals has changed into a multitude of tiny rhombic crystals. Rubbing the monoclinic sulfur with a few grains of powdered sulfur generally helps to start the transformation.

By rapidly cooling melted sulfur, we can make an amorphous, or non-crystalline, variety having rather peculiar properties.



EXPERIMENT 164—Plastic Sulfur.

Melt 5 or 6 measures of Sulfur (No. 1) in your spoon by heating rapidly over a flame. If you wish, you can simply re-melt the sulfur remaining in the spoon from the preceding experiment. When completely melted, pour the sulfur into a glass of cold water. Remove the congealed sulfur and you will find that it is soft and almost rubbery. After a couple of days it will harden and change back to the ordinary form. This is amorphous sulfur.

When a substance is divided into extremely small particles, too small to be seen with an ordinary microscope, yet larger than single molecules, we say that it is in the colloidal state. Such particles when suspended in a liquid have a number of interesting properties, the study of which is a part of colloid chemistry.

EXPERIMENT 165—Collodial Sulfur.

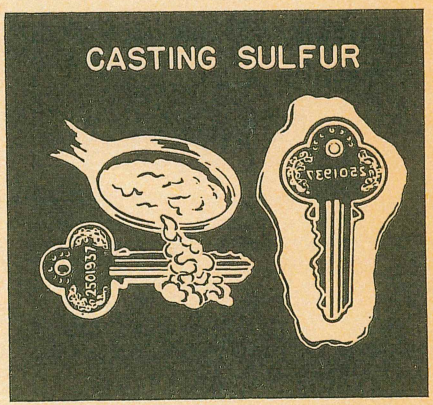
Dissolve 1 measure of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. In another test tube ¼ full of water place 1½ measures of Sodium Bisulfate (No. 7) and shake it until it is completely dissolved. Now pour the sodium bisulfate solution in the second tube into the sodium thiosulfate solution in the first tube. Close the tube with your thumb and mix the two liquids by inverting it.

Watch the liquid for the next few minutes and soon you will see it begin to become cloudy. When you hold it up to the light and look through it, it will appear yellowish or brown, but if you hold it against a dark background and let light shine into it from the side, it will look blue. This is because the light passing into the liquid is scattered by the tiny sulfur particles, and more of the blue rays of light are deflected than of the others. The color of tobacco smoke has a similar origin.

As the test tube stands, the particles grow larger, and presently the suspension will look nearly white unless the light passes directly through it, at which time it will look blue.

Particles as small as this pass right through the pores of ordinary filter paper. Try filtering some of your colloidal sulfur, and notice that the filtrate is about as cloudy as the original suspension.

Sulfur is sometimes used as a cement or as a casting medium, since it resists acids and is a good electrical insulator. It is poured while in a liquid form into a crevice or mold and allowed to solidify.



EXPERIMENT 166—Casting Sulfur.

Melt a small quantity of Sulfur (No. 1) in your spoon, being careful not to let it take fire. Pour the melted sulfur over a key or other small object. When it has solidified and the object is removed, you will find a perfect impression in the sulfur.

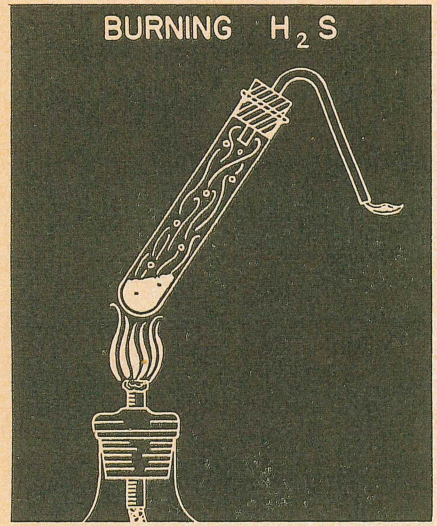
HYDROGEN SULFIDE

Hydrogen sulfide is a gas, each molecule of which is composed of two atoms of hydrogen and one atom of sulfur, H_2S. Everyone has smelled it, whether he realizes it or not, for it is the gas which is mainly responsible for the disagreeable odor of spoiled eggs and of the water in sulfur springs. Sewer gas also contains it. In spite of its foul odor, it is a very useful substance to the chemist, who uses it to precipitate the sulfides of the metals from solution in analytical work. The colors and characteristic reactions of the sulfides furnish a key to the metals which are present.

EXPERIMENT 167—Making Hydrogen Sulfide.

Cut from your candle a piece of paraffin about as big as the end of your little finger and place this together with 5 measures of Sulfur (No. 1) in a test tube. Heat the tube over a flame, and after a few minutes smell the hydrogen sulfide gas which is given off, first removing the test tube from the flame. Be careful not to smell very much of the gas, since it may give you a headache.

Paraffin is a hydrocarbon, that is, it is composed of hydrogen and carbon. When heated with sulfur, the hydrogen leaves the carbon and combines with the sulfur forming hydrogen sulfide.



EXPERIMENT 168—A Test for Hydrogen Sulfide.

Moisten a piece of Sulfide Test Paper (No. 31) and hold it in the mouth of the test tube in which hydrogen sulfide is being generated. The test paper will turn dark brown or black.

The test paper is impregnated with lead acetate, with which the hydrogen sulfide forms lead sulfide, a dark-colored, insoluble compound.

EXPERIMENT 169 — Burning Hydrogen Sulfide.

Fit the test tube in which hydrogen sulfide is being generated with the gas delivery tube and when the gas is coming off rapidly, apply a flame to the open end of the delivery tube, keeping your face at a little distance. The gas will burn with a blue flame.

When hydrogen sulfide burns in the air, the sulfur and hydrogen unite with oxygen, forming sulfur dioxide and water.

$$\underset{\text{(hydrogen sulfide)}}{2\,H_2S} + \underset{\text{(oxygen)}}{3\,O_2} \rightarrow \underset{\text{(sulfur dioxide)}}{2\,SO_2} + \underset{\text{(water)}}{2\,H_2O}$$

If, however, not enough air is present, only the hydrogen burns, and the sulfur is set free. Hold a cold piece of glass in the blue flame for a moment and notice the deposit of sulfur which forms on it.

EXPERIMENT 170—Testing for Hydrogen Sulfide in Illuminating Gas.

Test your gas supply for hydrogen sulfide by letting the gas from a jet flow against moistened Sulfide Test Paper (No. 31). Gas made from coal contains hydrogen sulfide, but this is supposed to be removed before the gas reaches the consumer.

EXPERIMENT 171—Hydrogen Sulfide from Boiling Cabbage.

The steam arising from cooking cabbage contains small amounts of hydrogen sulfide, formed from sulfur compounds in the cabbage. Hold a moistened strip of Sulfide Test Paper (No. 31) in the steam from a pot of boiling cabbage and after a minute or two it will darken, showing the presence of hydrogen sulfide.

EXPERIMENT 172—Tarnishing Metals with Hydrogen Sulfide.

Prepare some hydrogen sulfide by heating paraffin with sulfur as you did in one of the preceding experiments. Moisten a bright silver coin with water and hold it against the mouth of the test tube. After a minute or less, you will find a black stain of silver sulfide on the coin. Minute amounts of hydrogen sulfide in the air and other sulfur compounds in foods are the cause of the tarnishing of household silverware.

If instead of a silver coin, you use a penny, you will obtain a black stain of copper sulfide.

EXPERIMENT 173—Reaction of Hydrogen Sulfide with Chlorine.

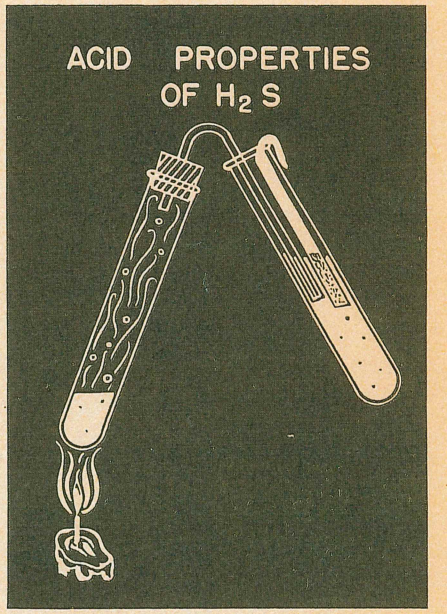
Generate hydrogen sulfide gas as above by heating with paraffin, and pass the gas through your gas delivery tube into a test tube ¼ full of water. Let the gas bubble through the water for a minute or two.

Now generate chlorine gas in another test tube by adding a few drops of water to a mixture of 2 measures of Bleaching Powder (No. 38) and 4 measures of Sodium Bisulfate (No. 7). Fit the delivery tube and stopper to this tube, and let the chlorine gas bubble through the hydrogen sulfide solution which you have just made. In a few moments the liquid will become cloudy and soon you will have a dense suspension of minute particles of sulfur—colloidal sulfur.

The chlorine takes the hydrogen away from the sulfur, forming hydrochloric acid, and the sulfur is left suspended in the liquid in the form of tiny solid particles.

$$\underset{\text{(hydrogen sulfide)}}{H_2S} + \underset{\text{(chlorine)}}{Cl_2} \rightarrow \underset{\text{(hydrochloric acid)}}{2\,HCl} + \underset{\text{(sulfur)}}{S}$$

When dissolved in water, hydrogen sulfide forms a small number of hydrogen ions, and hence it is a weak acid.



EXPERIMENT 174—Acid Properties of Hydrogen Sulfide.

Fill a clean test tube ¼ full of water and drop in a small piece of blue Litmus Paper (No. 18). Now generate hydrogen sulfide gas as before by heating sulfur with paraffin and pass it through your gas delivery tube into the water containing the litmus paper. After the gas has bubbled through the water for a few moments, the litmus paper will turn red, showing that the solution of hydrogen sulfide is acid.

Being an acid, hydrogen sulfide forms salts with bases. These salts are called sulfides. Most of the sulfides are decomposed by strong acids with the liberation of hydrogen sulfide. This is sometimes a convenient way of preparing the gas.

EXPERIMENT 175—Decomposing a Sulfide.

Prepare some iron sulfide by heating 2 measures of Sulfur (No. 1) and 2 measures of powdered iron in your spoon until they combine. Place 1 measure of this compound in a test tube. Add 3 measures of Sodium Bisulfate (No. 7) and about ½ inch of water. Heat over a flame until the liquid begins to boil.

Notice the bubbles which immediately begin to rise from the particles of iron sulfide. After a few moments you will be able to smell the hydrogen sulfide at the mouth of the test tube. Moisten a piece of the Sulfide Test Paper (No. 31) and hold it at the mouth of the test tube. The brown discoloration on the paper proves that hydrogen sulfide is being evolved.

Just as oxygen reacts directly with many metals to form oxides, so sulfur reacts forming sulfides. Many sulfides can also be formed by passing hydrogen sulfide gas into a solution of some salt of the metal. Most sulfides are insoluble in water and hence precipitate out from the solution. Some of the sulfides dissolve readily in dilute acids and therefore will not form if the solution is acid. Others are not affected and may be precipitated even from an acid solution. These differences are used in analyzing mixtures of metals.

EXPERIMENT 176—Direct Formation of Silver Sulfide.

Place ½ measure of Sulfur (No. 1) on a bright silver coin and wrap it in several thicknesses of paper. After a few hours you will find a black spot of silver sulfide on the coin where the sulfur was in contact with it.

$$\underset{\text{(silver)}}{2\,Ag} + \underset{\text{(sulfur)}}{S} \rightarrow \underset{\text{(silver sulfide)}}{Ag_2S}$$

A silver coin is turned black in a few hours by a paste of mustard and water, showing that mustard contains sulfur compounds. Eggs also contain sulfur, which explains why silverware turns black when used with eggs.

EXPERIMENT 177—Direct Union of Copper and Sulfur.

Place ¼ measure of Sulfur (No. 1) on a copper coin. Hold the coin with your test tube holder and heat it until the sulfur melts. Let it cool, scrape off the sulfur, and examine the black film of copper sulfide.

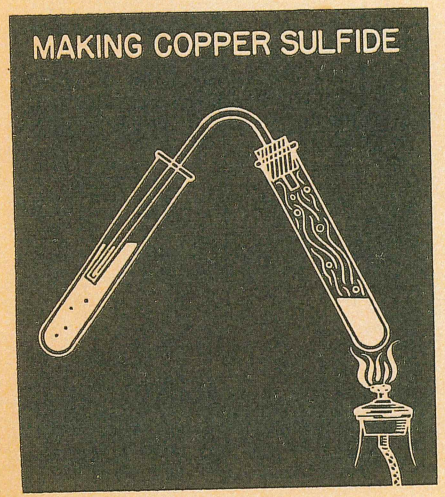
$$\underset{\text{(copper)}}{Cu} + S \rightarrow \underset{\text{(copper sulfide)}}{CuS}$$

EXPERIMENT 178—Proving That There is Sulfur in Rubber.

Rubber contains sulfur used in vulcanizing it. Wrap a rubber band around a silver coin, and after a few hours you will find black streaks of silver sulfide on it.

EXPERIMENT 179—Cobalt Sulfide.

Place 4 or 5 drops of Cobalt Chloride Solution (No. 27) and ½ measure of Sodium Carbonate (No. 4) in a test tube half full of water. A bluish precipitate of basic cobalt carbonate will form. Now pass hydrogen sulfide gas into the suspension and black cobalt sulfide will form from the particles of basic carbonate. The formula for cobalt sulfide is CoS.



EXPERIMENT 180—Copper Sulfide.

Place ¼ measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water and warm the mixture gently until the solids have completely dissolved and the liquid is clear. This is a solution of copper sulfate.

Now generate hydrogen sulfide by putting a piece of paraffin or candle wax the size of the end of your little finger into a test tube along with 5 measures of Sulfur (No. 1). Fit the gas delivery tube and stopper to the test tube and heat the mixture. Pass the evolved hydrogen sulfide into the copper sulfate solution, and black copper sulfide will precipitate. A small quantity of sulfuric acid is also formed. This is one of the sulfides which does not dissolve easily in acids.

$$\underset{\text{(copper sulfate)}}{CuSO_4} + \underset{\text{(hydrogen sulfide)}}{H_2S} \rightarrow \underset{\text{(copper sulfide)}}{CuS} + \underset{\text{(sulfuric acid)}}{H_2SO_4}$$

EXPERIMENT 181—The Reaction of Hydrogen Sulfide with a Ferric Salt.

Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. In another test tube, place a small piece of paraffin or candle wax with 5 measures of Sulfur (No. 1). Fit the gas delivery tube and stopper to this test tube and heat the mixture. As hydrogen sulfide gas is evolved, pass it into the test tube containing the ferric ammonium sulfate solution. Be very careful to keep the test tube in which hydrogen sulfide is being generated over the flame so that the ferric ammonium sulfate solution will not be sucked back into the hydrogen sulfide tube. After a short time, you will notice that a grayish white precipitate has formed in the tube which contained the ferric ammonium sulfate solution. This precipitate is sulfur formed by the oxidation of hydrogen sulfide by the ferric ions. The ferric ions are in turn reduced to ferrous ions. Were it not for this oxidation-reduction reaction in the case of the ferric salt, we would have obtained ferric sulfide, which is black.

EXPERIMENT 182—Zinc Sulfide

Place several small scraps of zinc metal in a test tube and add 3 measures of Sodium Bisulfate (No. 7) and ½ inch of water. As the zinc starts to dissolve, bubbles of hydrogen will come off. Heat the liquid to boiling to speed up the reaction, and let the test tube stand until no more bubbles come off. This may require 10 or 15 minutes. When the reaction is over, fill the test tube half full of water and bubble hydrogen sulfide gas through the solution. The white precipitate is zinc sulfide, ZnS.

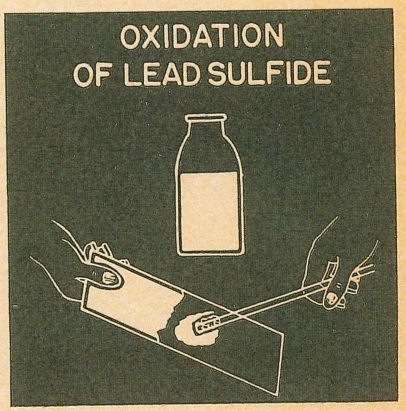
Zinc sulfide will not form if any acid is present, and therefore, it is important to wait until all the acid has been used by reacting with the zinc metal.

EXPERIMENT 183—Oxidation of Lead Sulfide.

Expose a moistened piece of Sulfide Test Paper (No. 31) to hydrogen sulfide gas. When it is thoroughly darkened, wet it with ordinary hydrogen peroxide, and it will turn white again. The hydrogen peroxide gives up oxygen to the black lead sulfide, converting it into white lead sulfate.

$$\underset{\text{(lead sulfide)}}{PbS} + \underset{\text{(hydrogen peroxide)}}{4\,H_2O_2} \rightarrow \underset{\text{(lead sulfate)}}{PbSO_4} + \underset{\text{(water)}}{4\,H_2O}$$

Many white paints contain lead carbonate, and traces of hydrogen sulfide in the air react with this compound to form lead sulfide, thus darkening the paint. Many famous paintings have suffered this fate, but fortunately the treatment with hydrogen peroxide restores their original appearance. For chemical laboratories, where hydrogen sulfide is common, zinc oxide is used as a white pigment in paint instead of the lead compound.



Many organic compounds contain sulfur in combination with carbon, hydrogen and other elements. Many of these have very unpleasant odors, as you will admit after performing the following experiment.

EXPERIMENT 184—Organic Sulfur Compounds.

Place 1 measure of Sulfur (No. 1) and 2 or 3 drops of glycerine in a dry test tube, and heat it slowly over a flame. As the glycerine begins to boil it will become reddish, and after a minute it will appear cloudy and have a red-brown color. Remove the test tube from the flame and let it cool slightly. At the mouth of the tube you will probably be able to recognize the odor of your friend hydrogen sulfide, tube you will probably be able to recognize the odor of your old friend hydrogen sulfide, but there will be another odor also, which may remind you of burning rubber. This is due to a mixture of organic sulfides and to a compound called allyl mercaptan.

SULFUR DIOXIDE

When sulfur burns in the air, the colorless gas, sulfur dioxide is produced. It has a sharp, choking odor, and is a heavy gas, being about twice as dense as air. Like chlorine, it is easily liquefied. Tremendous quantities are made by burning sulfur for use in the manufacture of sulfuric acid. Other important uses are for the preparation of sulfites for the paper industry, for bleaching various materials, and in numerous chemical manufacturing processes. Some of the household mechanical refrigerators use liquid sulfur dioxide in the same way as commercial plants use liquid ammonia.

EXPERIMENT 185—Making Sulfur Dioxide.

Heat one measure of Sulfur (No. 1) in your spoon over a flame until it melts and takes fire. Smell cautiously above the burning sulfur. Once you smell it, you will never forget the odor of sulfur dioxide.

$$S + O_2 \rightarrow SO_2$$

EXPERIMENT 186—Removing Fruit Stains with Sulfur Dioxide.

Stains caused by fruit juices are quickly bleached by sulfur dioxide. If you get some fruit stains on your hands, wet them with water and hold them over a little sulfur to which you have set fire. Fruit stains are also removed from fabrics by ths simple treatment. Wash your hands with soap and water as soon as the stain is bleached.

SULFURIC ACID AND THE SULFATES

In addition to sulfur dioxide, sulfur can form another oxide, containing an extra atom of oxygen. Its formula is SO_3, and its name is sulfur trioxide. When it reacts with water, it too forms an acid:

$$SO_3 + H_2O \rightarrow H_2SO_4$$
(sulfuric acid)

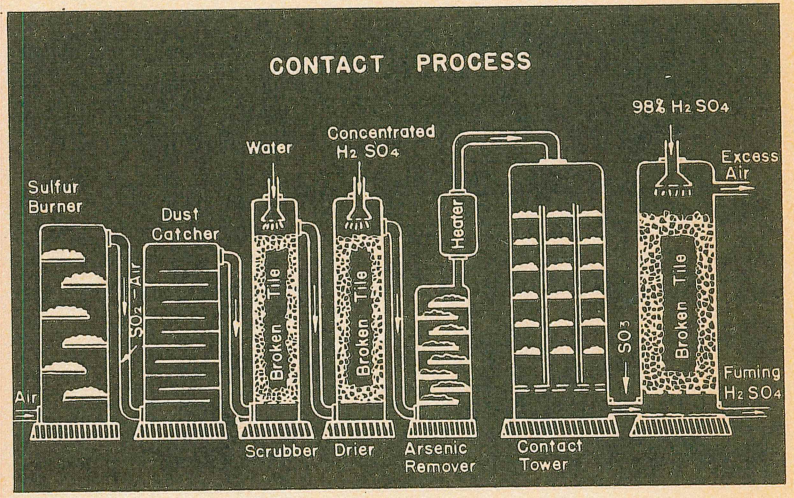
It is a general rule that when a non-metallic element forms more than one oxide, the one with the most oxygen in it gives the strongest acid, and therefore sulfuric acid is much stronger than sulfurous acid, which is formed when sulfur dioxide, SO_2, is dissolved in water.

Pure sulfuric acid is a heavy, colorless, oily liquid. It has a powerful attraction for water, and chars and destroys organic matter such as sugar, wood and paper. When diluted with water it readily dissolves many metals, giving off hydrogen and forming salts known as sulfates:

$$Fe + H_2SO_4 \rightarrow H_2\uparrow + FeSO_4$$
(iron) (sulfuric acid) (hydrogen) (ferrous sulfate)

Sulfuric acid is the cheapest of the strong acids and the most useful. There are no industries which do not use it directly or indirectly. More than a million tons a year are used in making fertilizers. The petroleum industry uses tremendous quantities in refining oils and gasoline, but the newer processes of solvent refining are beginning to drive it out of this field. The iron, steel, and other metallurgical industries would be seriously crippled without sulfuric acid. It is a necessity also in making a great variety of chemicals, in making explosives and various paints and pigments, and in the manufacture of rayon and cellophane. The liquid in the ordinary type of storage battery is moderately strong sulfuric acid.

There are two important methods for making sulfuric acid. In both of them, sulfur dioxide is the starting material. It is commonly made by burning sulfur, but sometimes it is recovered from waste smelter gases and sometimes it is made by roasting iron pyrites. In the chamber process, the sulfur dioxide is mixed in large lead chambers with oxygen and with oxides of nitrogen and steam. It is oxidized to sulfur trioxide, which combines with the steam forming sulfuric acid. Since the steam must be present in excess, the resulting acid is rather dilute.



When concentrated acid is desired, the contact process is generally used. Carefully purified sulfur dioxide is mixed with air and led over a catalyst. Originally, the catalyst was asbestos covered with finely divided platinum. The sulfur dioxide and oxygen combined on the surface of the tiny particles of platinum, forming sulfur trioxide. The platinum itself was not changed in the process and hence could be used over and over. Nowadays, a better catalyst, composed chiefly of vanadium pentoxide, is used.

$$2\ SO_2 + O_2 \text{ (catalyst)} \rightarrow 2\ SO_3$$

It is rather difficult to make sulfur trioxide combine directly with water to form sulfuric acid, and so the sulfur trioxide is absorbed first in previously made sulfuric acid, and then treated with water.

EXPERIMENT 187—Making Sulfuric Acid.

If you have a small amount of hydrogen peroxide, you can demonstrate the preparation of sulfuric acid. Obtain a half pint milk bottle or a similar wide mouth bottle and cut a small square of cardboard that you can place across the top to prevent the escape of gas. Now place 5 measures of Sulfur (No. 1) in your spoon, lay the spoon on a tin plate and light the sulfur. When the sulfur is burning well, hold the mouth of the bottle about an inch above the spoon so that the sulfur dioxide coming off from the burning sulfur will rise into the bottle. Hold the bottle in this position for several minutes until it is well filled with sulfur dioxide gas. Now quickly slip the piece of cardboard over the mouth of the bottle and turn it right side up, keeping the cardboard in place. Now add some hydrogen peroxide to the bottle (an amount about equal to a tablespoonful) and quickly replace the cardboard over the mouth of the bottle. Holding the cardboard in place with your

fingers, shake the bottle around so that the hydrogen peroxide will come in contact with the sulfur dioxide gas. Shake the bottle for a moment or two, then remove the cardboard and test the liquid in the bottle with a piece of blue Litmus Paper (No. 18). You will find that the litmus paper turns a bright red, indicating that you have an acid.

$$\underset{\text{(sulfur dioxide)}}{SO_2} + \underset{\text{(hydrogen peroxide)}}{H_2O_2} \rightarrow \underset{\text{(sulfuric acid)}}{H_2SO_4}$$

In a clean test tube ¼ full of water dissolve 1 measure of Strontium Chloride (No. 10) and add the few drops of sulfuric acid which you have made. A white precipitate of strontium sulfate forms, proving that the acid which you made is actually sulfuric.

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(strontium chloride)}}{SrCl_2} \rightarrow \underset{\text{(strontium sulfate)}}{SrSO_4} + \underset{\text{(hydrochloric acid)}}{2\,HCl}$$

EXPERIMENT 188—Precipitating Strontium Sulfate.

In a test tube ¼ full of water, dissolve 1 measure of Aluminum Sulfate (No. 12) and in another test tube ¼ full of water dissolve 1 measure of Strontium Chloride (No. 10). Pour the 2 solutions together and you will obtain a heavy white precipitate of strontium sulfate. This is the same thing chemically as the mineral celestite. The precipitation of strontium sulfate is the same reaction you used in the preceding experiment to test your sulfuric acid.

EXPERIMENT 189—Testing for Sulfuric Acid in a Storage Battery.

By means of a glass tube, remove a drop of the electrolyte from an automobile storage battery and drop it into a test tube ¼ full of water. Test this solution with a piece of blue Litmus Paper (No. 18) and you will find that the paper is immediately turned red, showing that the liquid is strongly acid.

In another test tube ¼ full of water, dissolve 1 measure of Strontium Chloride (No. 10). Pour a few drops of the diluted electrolyte from the first tube into the strontium chloride solution. The white precipitate of strontium sulfate proves that the electrolyte contains sulfate ions and since we have already shown it to be acid, we conclude that sulfuric acid must be present. This experiment will not work with the electrolyte from an alkali storage cell which contains sodium hydroxide solution instead of sulfuric acid.

SODIUM THIOSULFATE

Sodium thiosulfate is a salt of the very unstable thiosulfuric acid, $H_2S_2O_3$. Notice that this acid differs from sulfuric in having one more sulfur atom and one less oxygen atom in its molecule. It loses the extra sulfur atom very readily.

EXPERIMENT 190—Separating Sulfur from Thiosulfuric Acid.

Dissolve 2 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water and drop in 2 measures of Sodium Bisulfate (No. 7). Close the tube with your thumb and shake it well. Soon the solution will appear slightly turbid, and gradually a thick precipitate of sulfur will form. At the mouth of the tube you can smell sulfur dioxide.

Acids liberate thiosulfuric acid from sodium thiosulfate, and the unstable thiosulfuric acid then slowly decomposes into sulfurous acid and sulfur.

Although the salts of thiosulfuric acid, the thiosulfates, are much more stable than the free acid, they too break down when heated. If water is present, hydrogen sulfide is given off.

EXPERIMENT 191—Making Hydrogen Sulfide from Sodium Thiosulfate.

Place 1 measure of Sodium Thiosulfate (No. 16) in a dry test tube and heat it strongly over a flame. This salt contains water of hydration, part of which is first driven off, condensing in droplets on the upper part of the test tube. On continued heating, the remaining water will react with the thiosulfate, forming hydrogen sulfide. When the solid begins to turn dark, remove the test tube from the flame and smell carefully at its mouth for the familiar odor of hydrogen sulfide. Hold a moistened piece of Sulfide Test Paper (No. 31) at the mouth of the tube, and it will soon darken. Some free sulfur is also formed by the decomposition as you can see by examining the wall of the test tube just above the dark residue in the bottom.

EXPERIMENT 192—Formation and Decomposition of Ferric Thiosulfate.

Dissolve 2 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water, and in another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). Add a few drops of the ferric ammonium sulfate solution to the sodium thiosulfate solution, and you will obtain a dark red color. This is a solution of ferric thiosulfate. In a few seconds you will notice it beginning to fade, and after a minute or two the solution will be nearly colorless, since ferric thiosulfate decomposes rapidly forming ferrous tetrathionate, which has no color.

Since the thiosulfates, under certain conditions, may decompose producing hydrogen sulfide, the sulfides of some of the metals may be precipitated by means of sodium thiosulfate.

EXPERIMENT 193—Precipitating Cuprous Sulfide.

To a test tube ¼ full of water add ½ measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7). Warm over a flame until the solution becomes clear. It now contains copper sulfate. Heat it just to boiling and add 1 measure of Sodium Thiosulfate (No. 16). Boil for a few moments, taking care not to let the liquid boil over. First a yellowish color appears, then the solution becomes cloudy, and soon there is a thick brown precipitate. Let the tube stand for a few minutes, occasionally warming it again over the flame. The precipitate will slowly darken and finally a heavy black substance will settle to the bottom. This is cuprous sulfide, Cu_2S. The sulfide of copper which you precipitated with hydrogen sulfide in a previous experiment was cupric sulfide, CuS. The two sulfides resemble each other closely.

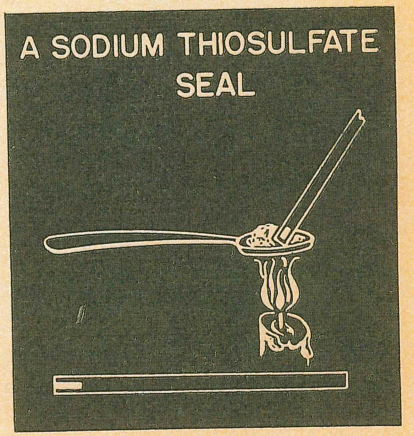
Sodium thiosulfate finds wide application in chemistry and in industry. The "hypo" which the photographer uses is nothing but sodium thiosulfate. It has the property of dissolving silver salts from negatives and prints. During the process of development, the silver salts which have been exposed to light are changed into finely divided metallic silver. After development, the films or prints are placed in the hypo "fixing" bath, which dissolves out the unchanged silver salts, but leaves behind the black deposit of metallic silver which forms the image. The fixing bath thus removes all the remaining light-sensitive compounds and makes the picture permanent.

In earlier sections we saw that sodium thiosulfate reacts with chlorine and iodine. This is the basis of its use as an "anti-chlor" to remove excess chlorine from textiles after bleaching. Its reaction with iodine is an important one in analytical chemistry. Other uses are in certain dyeing operations and in the sugar industry to prevent fermentation.

When heated, sodium thiosulfate melts easily in its water of crystallization and solidifies on cooling. Because of this behavior it has been used for sealing glass tubes containing explosives in underwater torpedoes.

EXPERIMENT 194 — A Sodium Thiosulfate Seal.

Melt 4 measures of Sodium Thiosulfate (No. 16) in your spoon by heating gently over a flame. Be careful not to heat it more strongly than necessary. Dip the end of a piece of glass tubing into the melt, tilt it a little and remove it in such a way that the end remains filled with the melted thiousulfate. When it cools you will find that the tube is tightly closed. Such seals, of course, are only of temporary value, since the solid salt is dissolved by water.



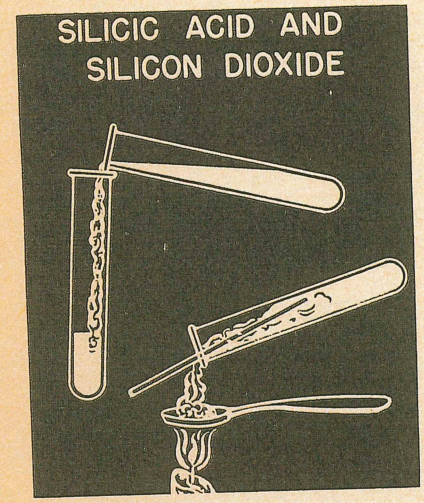
Silicon and Silicates

About a quarter of the earth's crust is silicon in combination with other elements. Quartz, sand, and sandstone are composed of more or less pure silicon dioxide, SiO_2, and many common minerals such as clay, feldspar, mica, soapstone, asbestos, and garnet consist of silicates of various metals. Free silicon is very similar to boron. It is not easy to prepare pure, but in the form of an alloy with iron it finds wide application in the iron and steel industry.

Silicon is a non-metallic element occuring in Group IVb of the Periodic Table. It exhibits a valence of 4.

The compounds of silicon are very numerous and some of them are of great importance. In addition to the natural compounds of silicon mentioned above, several artificial ones are worth mentioning. Silicon carbide, SiC, made from sand and coke in electric furnaces, is sold commercially as Carborundum for use in making grinding wheels. Sodium silicate is made by heating sand with sodium carbonate or sodium hydroxide. When dissolved in water it is called water glass and is used as an adhesive and in the manufacture of soaps, pottery glazes, and many other materials.

EXPERIMENT 195—Making Silicic Acid.



Put ½ inch of Sodium Silicate Solution (No. 19) in a test tube and add water until the tube is ¼ full. Shake the test tube to mix the liquids.

In another tube put 4 measures of Sodium Bisulfate (No. 7) and fill the tube ⅓ full of water. Shake the tube until the solid is completely dissolved.

Pour the sodium silicate solution into the sodium bisulfate solution and mix. A jelly-like precipitate will form, and in a few minutes all the liquid in the tube will become solid.

Save this silicic acid jelly for use in the next experiment.

EXPERIMENT 196—Making Silicon Dioxide.

Remove some of the silicic acid formed in the preceding experiment from the test tube by means of a stirring rod and put some of it in your spoon. Heat the spoon strongly in an alcohol lamp flame for 4 or 5 minutes, or until the precipitate is entirely dried and has been reduced to a white solid. This substance is silicon dioxide, the same thing chemically as pure sand or quartz, mixed with sodium sulfate, a by-product of the reaction in the preceding experiment.

EXPERIMENT 197—Soluble Glass (Sodium Silicate).

Paint a thin film of Sodium Silicate Solution (No. 19) on a sheet of paper and let it dry for 10 or 15 minutes. Notice the smooth, transparent, glass-like film which results.

Paste together two sheets of paper or two blocks of wood, using sodium silicate solution as the adhesive. You will find that it makes an exceptionally strong adhesive, and it is often used for this purpose.

Silicates of sodium and potassium are the only silicates which are readily soluble in water.

The most important artificial silicates are glass and portland cement. Glass is to be discussed in a later section. Portland cement is made by strongly heating a mixture of calcium carbonate (limestone) and a siliceous clay. The calcium unites chemically with the silica and alumina of the clay, producing a substance which when very finely ground and mixed with water, sets to a solid stone.

Clay is a combination of silica with aluminum oxide and water. It is the combined water which gives the property of plasticity, and makes it possible to mold it into pottery and bricks. When the molded article is "burned," that is, heated strongly, the combined water is driven off and the clay is made permanently hard.

Most metallic silicates are insoluble in water. They are formed as beautiful precipitates when a soluble silicate is added to a solution of the metallic salt.

EXPERIMENT 198—Aluminum Silicate.

Place 2 measures of Aluminum Sulfate (No. 12) in a test tube and fill the tube half full of water, shaking it to completely dissolve the solid. Now add 2 or 3 drops of Sodium Silicate Solution (No. 19) and notice the thick precipitate of aluminum silicate which is immediately formed.

EXPERIMENT 199—Cobalt Silicate.

Add 3 or 4 drops of Cobalt Chloride Solution (No. 27) to a test tube half full of water. Now add 2 or 3 drops of Sodium Silicate Solution (No. 19). In this case, a beautiful blue precipitate of cobalt silicate is formed. This is one of the most colorful of the entire series of silicates.

EXPERIMENT 200—Copper Silicate.

Place 3 measures of Sodium Bisulfate (No. 7) and 1 measure of Azurite (No. 39) in a test tube and fill the tube half full of water. Heat the liquid for a few moments until a clear blue solution of copper sulfate is obtained. Add to the copper sulfate solution 2 or 3 drops of Sodium Silicate Solution (No. 19), and examine the blue precipitate of copper silicate which is formed.

EXPERIMENT 201—Ferric Silicate.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water and add 2 or 3 drops of Sodium Silicate Solution (No. 19). A very pretty, reddish brown precipitate of ferric silicate will be formed.

EXPERIMENT 202—Magnesium Silicate.

Dissolve 2 measures of magnesium sulfate (Epsom salts) in a test tube half full of water. Add to this solution 2 or 3 drops of Sodium Silicate Solution (No. 19). A white precipitate of magnesium silicate will be formed.

EXPERIMENT 203—Strontium Silicate.

Dissolve 2 measures of Strontium Chloride (No. 10) in a test tube half full of water and add 2 or 3 drops of Sodium Silicate Solution (No. 19). A bulky, white precipitate will form and upon shaking the test tube the precipitate will fill the whole tube.

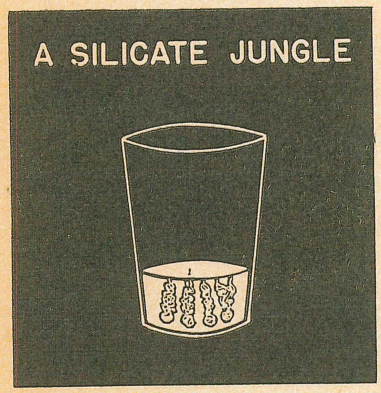
EXPERIMENT 204—Zinc Silicate.

Place a small piece of zinc metal and 2 measures of Sodium Bisulfate (No. 7) in a test tube. Fill the tube ¼ full of water, and warm the solution for a few moments until some of the zinc dissolves. Now hold the tube in a glass of cold water until it becomes cool. Add to the solution of zinc sulfate thus formed 2 or 3 drops of Sodium Silicate Solution (No. 19). A thick precipitate of zinc silicate will be formed.

EXPERIMENT 205—Reaction of Ammonium Chloride with Sodium Silicate.

We have been making silicates by adding sodium silicate solution to solutions of other salts, and you might ask now whether we could make ammonium silicate by adding sodium silicate solution to an ammonium salt, for example, ammonium chloride. The answer is that if ammonium silicate is formed, it decomposes at once, since its acid, silicic acid and its base, ammonium hydroxide, are both so weak that they cannot hold each other. Furthermore, ammonium hydroxide decomposes very easily into ammonia gas and water. Another way of stating the matter is to say that ammonium silicate is completely hydrolyzed by water, forming silicic acid and ammonium hydroxide. The ammonium hydroxide then proceeds to decompose.

Dissolve 2 measures of Ammonium Chloride (No. 9) in a test tube ¼ full of water and add 2 or 3 drops of Sodium Silicate Solution (No. 19). Close the mouth of the tube with your thumb at once and shake vigorously. The contents of the tube turn to a white slush or jelly of silicic acid, and if you smell at the mouth of the tube you can detect the ammonia formed in the reaction.



EXPERIMENT 206 — A Silicate Jungle.

Place 2 spoonfuls of Sodium Silicate Solution (No. 19) in a tumbler and fill it ¼ full of water. Mix the two liquids and drop in crystals of some of the following salts: Ferric Ammonium Sulfate (No. 21), Aluminum Sulfate (No. 12), Strontium Chloride (No. 10) and magnesium sulfate (Epsom salts). Let the glass stand quietly for a few hours and observe the plant-like growth which springs up from the crystals. This is formed of silicate membranes.

Two Other Non-Metallic Elements

There are two other non-metallic elements which should be mentioned at this point. They are boron and phosphorus.

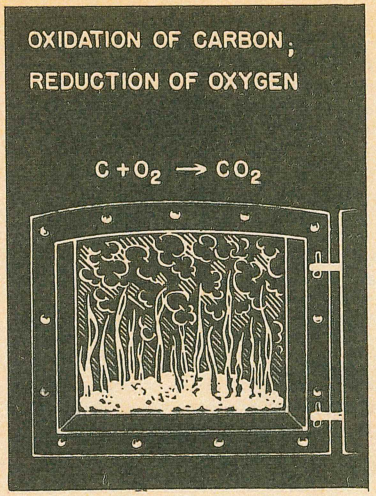
Boron occurs in Group IIIa of the Periodic Table and exhibits a valence of 3. The most important compound of boron is sodium borate or borax, $Na_2B_4O_7.10H_2O$. The element boron is exceedingly hard, being second only to the diamond.

Phosphorus occurs in Group Vb of the Periodic Table. Like nitrogen, which it follows in Group Vb, phosphorus exhibits five states of valence, from 1 to 5. Phosphorus precedes sulfur in period 3 of the Periodic Table and like sulfur, occurs in two physical forms, namely, white phosphorus and red phosphorus.

Phosphorus is a very interesting element and forms a large number of useful compounds, both with metals and with other non-metals. Phosphorus is found in nature in the form of phosphates, the most abundant being the mineral phosphorite, a calcium phosphate, $Ca_3(PO_4)_2$. Calcium phosphate is the basis of most of our compounds of phosphorus, including superphosphate, a calcium hydrogen phosphate, $Ca(H_2PO_4)_2.H_2O$, which is an important ingredient of fertilizer. Trisodium phosphate, known as TSP, is used extensively in soap powders and as an aid in cleansing.

Oxidation and Reduction

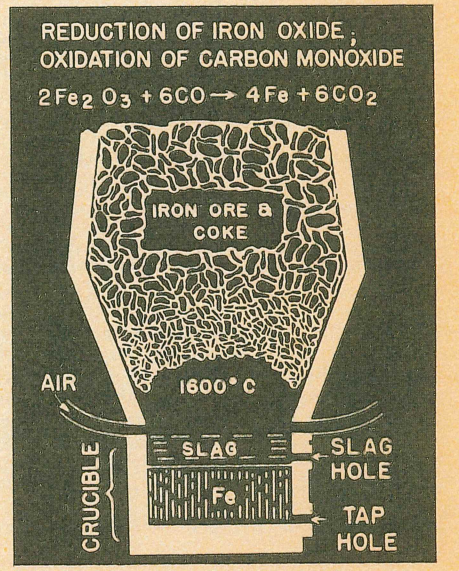
Many of the most important chemical processes of industry and of daily life depend upon reactions in which the valence of elements changes. Such reactions are called oxidation-reduction reactions. The burning of fuels,



the smelting of metals, the operation of dry-cells and of storage batteries, the rusting and tarnishing of metals, and the growth of plants and animals are intimately connected with oxidation-reduction reactions. When iron is not combined with any other element, you can say that its valence is zero because it is not using any of its capacity for combining. Suppose now that one allows the iron to react with oxygen, forming ferrous oxide, FeO:

$$2\ Fe. + O_2 \rightarrow 2\ FeO$$

The valence of iron in this compound is plus two, so the iron has increased two in valence. An increase in valence is called oxidation. Now let's see what has happened to the oxygen. The free oxygen, before the reaction, had a valence of zero because it was not combined with another element. After the reaction its valence is minus two. Its valence then went from zero to two below zero, that is, it decreased two. A decrease in valence is called reduction. In this reaction the iron was oxidized by the oxygen, and the oxygen was reduced by the iron. Whenever an atom is oxidized, some other atom must always be reduced.



We have already met examples of oxidation and reduction reactions. When we burned sulfur, we oxidized it, and simultaneously an equivalent amount of oxygen from the air was reduced by combining with the sulfur.

When we heated iron with sulfur, making iron sulfide, the reaction was very similar to the burning of iron in air. The only difference was that the iron combined with sulfur instead of with oxygen. Chemists extend the word oxidation to include the combination of non-metals with metals, regardless of whether there is any oxygen in the reaction or not. A similar statement goes for reduction also.

Now there is still more to this matter of oxidation and reduction. We find that many of the elements can combine with various fixed amounts of

other elements. An atom of iron, for example, can combine with two atoms of chlorine, or with three atoms. We say, therefore, that it may have a combining capacity of either two or three (these numbers are considered as plus, or positive, since iron forms positive ions), while chlorine has a combining capacity of minus one.

When we liberate iodine from sodium iodide solution by passing in chlorine:

$$2\,NaI + Cl_2 \rightarrow 2\,NaCl + I_2$$

we change its combining power from minus one to zero. This is, in a certain sense, an increase in combining power, and hence the iodine is oxidized. The chlorine at the same time changed to sodium chloride, in which its combining power is minus one. Its change, then, is from zero (free chlorine) to minus one, and thus it is reduced. Oxidation and reduction always go together.

All branches of chemistry must deal with oxidation and reduction reactions—metallurgy is based upon them, electrochemistry and phyiscal chemistry make use of them, organic chemists continually discuss them, and they lie at the very root of bio-chemistry, the chemistry of life.

EXPERIMENT 207—Formation of Iron Oxide.

Hold a small bright iron nail with a pair of pliers and heat it in the flame of the alcohol lamp for a few minutes. Plunge it into a glass of water and when it is cooled remove it and notice the changed appearance. The black color is due to iron oxide.

EXPERIMENT 208—Formation of Copper Oxide.

Hold a piece of bare copper wire with a pair of pliers and heat it in the flame of the alcohol lamp. The copper will impart a beautiful green coloration to the flame and when the copper is removed and plunged into a tumbler of water a black scaly deposit of copper oxide will form.

EXPERIMENT 209—Formation of Zinc Oxide.

Cut off a piece of zinc metal about the size of your fingernail and scratch it with a knife until it is shiny. Drop this into a test tube and heat it strongly over the alcohol lamp flame. After a few minutes you will notice a white deposit of zinc oxide appearing on the metal.

An interesting example of the oxidation of metals is the process of cutting with the oxy-acetylene flame. Oxygen and acetylene gases are mixed and the gas is lighted. This makes a flame so hot that it may be used in cutting through iron and steel of almost any thickness. The metal is heated to its kindling temperature and burned away.

Oxidation is not always accompanied by light or by noticeable heat. For example, when iron rusts it combines slowly with oxygen. When wood decays the products are nearly the same as those formed when it burns. Just as much heat is given off when iron rusts as when it is burned at a high temperature, but in the rusting the action requires such a very long time and the heat is given off so slowly that it is not noticeable, as it has no chance to accumulate.

EXPERIMENT 210—The Slow Oxidation or Rusting of Iron.

Put 2 or 3 measures of powdered iron in a tumbler and moisten it with a very few drops of water. Set the tumbler aside for a few hours and then notice the rust which is beginning to form. The iron rust which is formed in this experiment is very similar to the iron oxide which is formed when iron is burned at a high temperature.

An excellent example of reduction is the action of thermit, a mixture of powdered aluminum metal and iron oxide. Here the aluminum metal is used to combine with oxygen from the iron oxide just as carbon is used in the blast furnace.

In making welds, thermit is placed in a crucible having a small tap hole in the bottom and the work to be welded is contained in a mold immediately below the tap hole. The thermit is ignited and a violent reaction immediately takes place. The aluminum metal takes the oxygen away from the iron oxide forming aluminum oxide and metallic iron.

$$\underset{\text{(aluminum)}}{2\,Al} + \underset{\text{(iron oxide)}}{Fe_2O_3} \rightarrow \underset{\text{(aluminum oxide)}}{Al_2O_3} + \underset{\text{(iron)}}{2\,Fe}$$

The heating resulting from the union of the aluminum and the oxygen is intense enough to melt the iron. The tap hole in the bottom of the crucible is then opened and the molten iron runs into the mold, welding the piece which is held there.

In a broader sense the terms "oxidation" and "reduction" refer to reactions which are equivalent to the addition or removal of oxygen, even if no oxygen is actually present.

Metals

Of all the chemical elements, more than three-fourths are classed as metallic,which means that they can be made to combine with oxygen forming oxides which have basic properties and which unite wit hacids to form salts.

The most striking physical characteristics of metals are their luster, their ability to conduct electric current and to conduct heat. Everyone is familiar with the common metals, such as iron, copper, lead and tin and with their characteristics of weight, strength and metallic luster. There are, however, some metals which lack most of these properties and the real distinction between metals and non-metals is based upon chemical propreties.

Very few metals are found in nature in the pure state. Mostly they occur as ores which consist of various compounds of the metals. Iron, for example, occurs in the form of iron oxide, iron sulfide and iron carbonate. Copper occurs mostly as copper sulfide and copper carbonate. Lead occurs principally as lead sulfide. Gold, on the other hand, occurs in the free state mixed with various other minerals.

In order to obtain pure metals from the ores in which they occur in nature, it is necessary to devise certain methods of liberating the metals from their combinations. This is done mostly through oxidation and reduction reactions. In many cases the ore is first roasted or oxidized to remove impurities. The resulting oxide of the metal is then subjected to a process of reduction and the metal is obtained in the free state. Further operations may then be required to purify it. The art and science of extracting metals from their ores is called metallurgy.

ALLOYS

In many cases molten metals dissolve in one another freely. The mixtures, or compounds, resulting are called alloys. Many alloys have entirely different properties from those of the metals which are used in making them.

An amalgam is an alloy of mercury with some other metal. Copper forms alloys with many other metals. Brass contains various properties of copper and zinc. A yellow brass, for example, may run about 30% zinc and 70% copper. A typical bronze contains about 5% of tin, 15% of zinc, a small

quantity of lead and about 75% of copper. Gunmetal and bell metal are both mixtures of copper and tin. German silver is a mixture of copper, zinc and nickel.

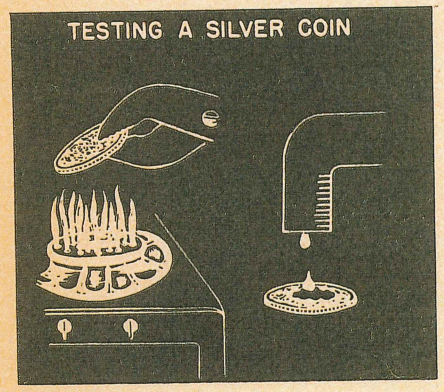
Pure iron is a relatively soft metal and could not be employed for the uses that are demanded of wrought iron and steel were it not mixed with other materials to form tough, durable alloys called steel.

Most coins are alloys of various metals, as the pure metal is in most cases too soft to stand the wear to which they are subjected. Our 5 cent piece, for example, is an alloy containing 75 per cent copper and 25 per cent nickel. Our silver coins contain 90 per cent of silver and 10 per cent of copper.

EXPERIMENT 211—Testing a Penny for Copper.

Heat a copper cent in a gas flame or over a coal fire until it becomes red hot. Notice the very pretty colors which play over the surface of the coin while you are heating it. Now let a few drops of water fall on the coin and when it becomes cool notice the formation of black copper oxide which covers it.

EXPERIMENT 212—Testing a Silver Coin to Show the Presence of Copper.



Repeat the preceding experiment with a silver coin, and notice that a somewhat lighter coat of black copper oxide is obtained in this case. Pure silver will not oxidize in the air even when heated. The black coating is copper oxide, and the presence of this oxide on the silver proves that coins are not pure silver.

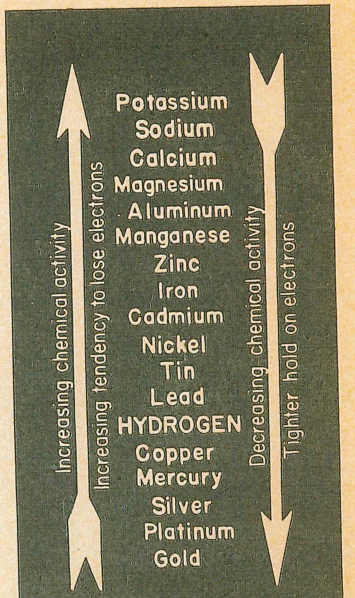
EXPERIMENT 213—Testing a Nickel Coin to Show that It Contains Copper.

Heat a 5-cent piece in a gas flame or over a coal fire until it becomes red hot, and then let a few drops of water fall on it. You will notice the same formation of black copper oxide which was obtained when you heated the cent.

The melting point of an alloy is often lower than the melting point of either of the metals of which it is composed. Alloys with low melting points are frequently used in automatic sprinkler systems and on fire doors so that if a fire starts and the metal gets hot it melts and releases the water from the sprinkler system or allows the fire doors to close automatically. One of the most common alloys with a low melting point is known as Wood's metal. It consists of 4 parts of bismuth, 2 parts of lead, 1 part of tin and 1 part of cadmium, and melts at a temperature much lower than the boiling point of water.

ACTIVITY SERIES OF THE METALS

In the discussion of metallic elements in the following pages, you will find certain metals referred to as active and others as inactive. The relative chemical activity of a metal depends upon the ease with which electrons are released from an atom of the metal. Since the release of electrons by the atom of a metallic element is just what occurs when the metal reacts with oxygen to form an oxide, the term "oxidation" has been broadened to cover all similar reactions. A metal, then, which loses electrons easily will combine readily with oxygen and also with other non-metallic elements. The accompanying table shows the relative activity of the metallic elements, starting with the most active and ending with the least active. Hydrogen is included in the table because, like the typical metals, it combines with oxygen and can undergo many other oxidation reactions by losing electrons. Any metal which is more active than hydrogen, that is, which stands above hydrogen in the table, can displace hydrogen in the form of gas from an acid solution. In previous experiments you used iron, aluminum, and zinc to generate hydrogen by reaction with acid, but copper and silver would not work, since they are less active than hydrogen, standing below it in the table.



ACTIVITY SERIES—
The best known metallic elements arranged in the order of their activity

THE ALKALI METALS

Group Ia of the Periodic Table is known as the alkali metals and consists of lithium, sodium, potassium, rubidium and cesium. The metals in this group all have the valence of one. Since these metals readily release an electron, they are chemically quite active and readily form a large number of compounds. The ions formed by the alkali metals are colorless and with but few exceptions their compounds are white and are readily soluble in water.

SODIUM

Most important among the alkali metals is sodium. In the metallic form, sodium is a silvery, soft, waxy substance which combines directly with water forming sodium hydroxide.

The chief source of sodium and its compounds is, of course, sodium chloride (common salt). All animal life requires sodium chloride and this salt is the principal raw material for practically all sodium compounds, among the most important of which are sodium hydroxide, sodium carbonate, sodium bicarbonate and sodium sulfate.

Sodium hydroxide is a strong base and because of its corrosive action, it is said to be "caustic." Sodium hydroxide is often called caustic soda. It is readily soluble in water, forming a strongly alkaline solution.

EXPERIMENT 214—Making Sodium Hydroxide.

Put 2 measures of Calcium Oxide (No. 20) and 2 measures of Sodium Carbonate (No. 4) in a test tube. Fill the tube ¼ full of water and boil the liquid for a minute or two, being careful that it does not boil over. Keep your face away from the tube while heating the liquid.

Allow the solution to cool and the solid matter to settle, then put a drop of the clear liquid on the end of your finger and rub it around, noticing the soapy feel. This is characteristic of strong bases like sodium hydroxide.

Sodium carbonate, sometimes called soda ash, is an important compound of sodium and is quite widely used in industrial processes, as well as in the home as washing soda.

EXPERIMENT 215—Making Sodium Carbonate from Sodium Bicarbonate.

Dissolve 1 measure of sodium bicarbonate (baking soda) in a test tube half full of water. Add 1 drop of Phenolphthalein Solution (No. 22) to the sodium bicarbonate solution and notice that a very faint pink color results.

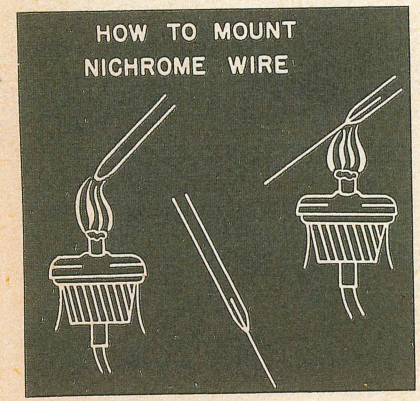
Now carefully boil the solution for a few moments, and you will find that the pink color becomes a deep red color.

Sodium bicarbonate is a very mild alkali, but when heated it gives off carbon dioxide and changes into sodium carbonate which is a stronger alkali. Thus, when you heated the sodium bicarbonate solution you changed it into sodium carbonate solution.

Many metals and their salts possess the very interesting property of imparting characteristic colorations to a flame, such as that of an alcohol lamp or bunsen burner. This property makes it possible to identify certain metals in a chemical compound. To make these flame tests you will need a piece of Nichrome Wire (No. 58) fused into a glass tube handle.

EXPERIMENT 216—How to Mount your Nichrome Wire.

Heat the end of a piece of glass tubing, approximately six inches long, in the flame of the alcohol lamp, turning the tube constantly so that it is heated evenly. When it becomes soft, you will notice that the end of the tube is gradually closing. At this point remove the tubing from the flame and insert one end of the Nichrome Wire (No. 58) into the small opening for about one-half inch. Again heat the end of the tube until the wire is fused into the glass.



Allow the tube to cool and then fire polish the opposite end of the tube by holding it in the flame of the alcohol lamp until the rough edges are rounded.

EXPERIMENT 217—Flame Color Test for Sodium.

Moisten your nichrome wire and hold it in the flame of the alcohol lamp until the flame is no longer colored by the wire. Now again moisten the nichrome wire and touch it to some sodium chloride (common table salt), hold the wire in the alcohol lamp flame and notice the bright yellow color that is imparted to the flame. This is characteristic of sodium and sodium compounds. The sodium flame is very persistent and it may take quite a little time to burn all of the sodium off your nichrome wire. Dipping the wire in clear water occasionally and then holding it again in the flame will help to get rid of this color.

Try this test with other sodium compounds in your outfit, such as sodium bisulfate, sodium thiosulfate and sodium carbonate.

POTASSIUM

Potassium is very similar, both physically and chemically, to sodium. The pure metal is slightly more active than sodium metal and when a piece of potassium is dropped into water the reaction is so vigorous that the hydrogen liberated from the water catches fire and burns.

Potassium is not as plentiful as sodium although potassium compounds are very widely distributed throughout the earth. All growing plants need potassium and this element must, therefore, be in the soil in such a form that the plants can absorb it. Most mixed fertilizers contain potassium in the form of one of its soluble salts.

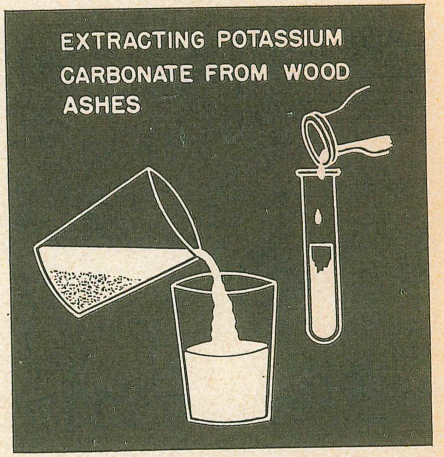
Potassium salts are obtained from many different sources, principally from deposits left by the drying up of lakes or inland seas. Seaweed contains a large per cent of potassium salts and potassium compounds are also obtained as a by-product from cement kilns. Formerly, the principal source of potassium was wood ashes.

EXPERIMENT 218—Extracting Potassium Carbonate from Wood Ash.

Put 2 or 3 spoonfuls of fresh wood ashes in a tumbler and fill it ½ full of water. Stir for a few minutes and allow the solid matter to settle to the bottom. Pour off some of the clear solution into another glass.

Test a small portion of this liquid in a test tube by adding a few drops of Phenolphthalein Solution (No. 22) and also by dipping into it a piece of red Litmus Paper (No. 18). You will find that it gives an alkaline reaction because of the potassium carbonate dissolved out of the wood ashes.

The product obtained by evaporating this solution of impure potassium carbonate is known as "pearl ash" and was formerly an important industrial product.



EXPERIMENT 219—Flame Color Test for Potassium.

Clean your flame test wire by heating it in the flame of the alcohol lamp until it no longer gives a color to the flame. Then dip it into the solution obtained from the wood ashes in the preceding experiment and heat it again in the flame. Look at the flame through a piece of blue glass and you will see the characteristic violet color given to a flame by potassium.

THE ALKALINE EARTH METALS

The elements of Group IIa in the Periodic Table are called alkaline earth metals because their oxides were called alkaline earths in the early days of chemistry. The metals comprising this group are beryllium, magnesium, calcium, strontium, barium, and radium. All of these metals have a valence of two.

As a group, the alkaline earth metals are second in activity only to the alkali metals. With the exception of beryllium, these metals are fairly soft.

The only metal of the alkaline earth group which is very useful in the metallic form is magnesium. This metallic element is very strong in pro-

portion to its weight and therefore finds considerable commercial use where strength and extreme lightness are required.

CALCIUM

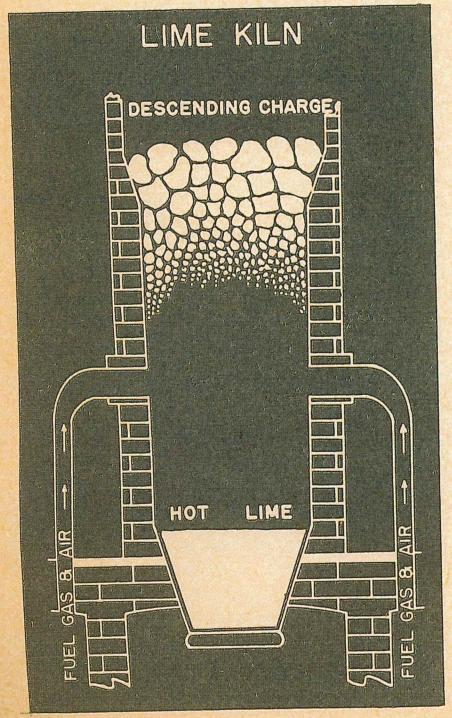
Calcium, which is a silvery white metal, decomposes water, forming calcium hydroxide and liberating hydrogen:

$$Ca + 2\,H_2O \rightarrow Ca(OH)_2 + H_2\uparrow$$

It is one of the most abundant elements and in combination with other elements forms nearly 3½% of the earth's crust. In combination with carbon and oxygen it forms calcium carbonate (limestone), with phosphorus and oxygen it occurs in phosphate rock, and with sulfur and oxygen it forms gypsum. Pearls are largely calcium carbonate, the same thing chemically as limestone.

A wide variety of calcium compounds are used in industry. Calcium oxide, or lime, is used for making mortar and plaster, and a complex compound of calcium with chlorine and oxygen is called bleaching powder, and is used in bleaching and as a disinfectant. Portland cement is made by heating a mixture of limestone and clay in kilns to an extremely high temperature and then pulverizing to a fine powder. The result is a mixture containing chiefly various calcium silicates and calcium aluminates.

Calcium oxide or lime is made by heating calcium carbonate.



EXPERIMENT 220—How to Make Lime.

If you can obtain some small pieces of limestone or marble, or even oyster shells, you can very easily make a small quantity of lime. Place 2 or 3 teaspoonfuls of the limestone or other material in the lid of a baking powder can or similar container. The limestone should be broken into rather small lumps.

The stone must now be heated to a very high temperature so it will be necessary to set the lid on a hot stove. Heat the limestone until it becames a bright red or nearly white color, that is, red hot, and then remove it from the fire and allow it to cool. You will find that it is different from the original limestone, in fact it is changed into calcium oxide or lime.

EXPERIMENT 221—Showing How Lime Expands When It is Slaked or Hydrated.

Fill a very small bottle or vial with finely ground lime. Pour a little water into the bottle and set the bottle aside to see what happens. As the lime slakes, it expands and occupies a greater volume. Since you have the lime confined in the bottle it will either have to break the bottle or come out of the mouth as it slakes. Be extremely careful when performing this experiment

EXPERIMENT 222—Flame Color Test for Calcium.

Moisten your Nichrome Wire (No. 58) and touch some Bleaching Powder (No. 38). This compound contains some calcium hypochlorite and when you hold the wire in the alcohol lamp flame, you will notice the yellowish red color, which is characteristic of calcium. The flame color test for calcium will not work with calcium oxide since the volatilization or boiling point of lime is 2850° C. This temperature cannot be attained in the alcohol lamp or bunsen burner flame.

MAGNESIUM

Magnesium is a silvery white metal which oxidizes to some extent in the air. Its compounds are widely distributed in the earth. Commercially, magnesium compounds are obtained from salt beds and natural brines. It is purified to the metallic form by electrochemical means. Powdered magnesium metal and magnesium ribbon burn brightly in the air and are used in fireworks. Magnesium is also used in flashbulbs for indoor photography. Magnesium is a constituent of duralumin and other light alloys which are important in airplane construction. Magnesium oxide, MgO, is used in lining high temperature furnaces and in making sorrel cement for stucco and floors. Milk of magnesia, used in medicine for neutralizing acidity, is a suspension of magnesium hydroxide in water. Epsom salts consist of magnesium sulfate with water of hydration, $MgSO_4.7H_2O$.

STRONTIUM

The compounds of strontium resemble those of calcium, both in physical properties and in chemical behavior. Most of our strontium is obtained from the mineral celestite, strontium sulfate. The carbonate of strontium, strontianite, is also found in nature. The pure metal is obtained by electrolysis of molten strontium chloride.

Strontium hydroxide is used in sugar refining, and strontium nitrate and strontium chloride are used in fireworks and in the manufacture of red flares and signals.

EXPERIMENT 223—Flame Color Test for Strontium.

Moisten your nichrome wire and touch it to some Strontium Chloride (No. 10). Hold the wire in the alcohol lamp flame and notice the beautiful red coloration that is imparted to the flame. Strontium salts give this characteristic red color to flame.

COPPER, SILVER AND GOLD

The metals copper, silver and gold are placed in Group Ib of the Periodic Table. Each of these metals may have a valence of 1, 2, or 3. Copper usually exhibits a valence of 1 or 2. Silver usually has a valence of 1 and gold usually exhibits a valence of 1 or 3. They are generally encountered with a valence of one. All three of these metals are difficult to oxidize and they are, therefore, frequently found in the metallic state in nature and to a great degree resist corrosion. Copper is the most easily oxidized of the three metals. Silver is more difficult and gold cannot be oxidized by direct means. Due to their resistance to oxidation and their relative chemical inactivity, silver and gold are known as the "noble metals."

COPPER

Copper metal has been known and used since prehistoric times, principally because it often occurs in nature as the metallic element and its physical characteristics allowed it to be formed easily into weapons and utensils. Among the more important ores of copper are various sulfides, such as chalcopyrite and chalcocite, an oxide, suprite and a basic carbonate, malachite. The copper is separated from its ores by roasting followed by

smelting, and the impure metal which is obtained is refined by the electrolytic process.

Copper is one of the few metals which has a distinguishing color. Owing to its property of conducting electricity, large quantities of this metal are used in the electrical industries. Copper is also used in equipment for chemical plants, in the building industries and in making electrotypes for printing. We have already mentioned some of the alloys of copper, such as brass and bronze, which are of great importance.

Since copper exhibits two valences, it forms two oxides, cuprous oxide, Cu_2O (valence 1) and cupric oxide, CuO (valence 2). Corresponding to these two oxides, there are two series of copper salts, the cuprous and cupric. Cupric salts are of great importance and cupric sulfate, $CuSO_4$, sometimes called bluestone, or blue vitriol, has use in many industrial and commercial applications. Dissolved in water and mixed with lime, cupric sulfate forms Bordeaux mixture, an important spray material for controlling fungus diseases of plants.

EXPERIMENT 224—Flame Color Test for Copper.

Moisten your nichrome wire and heat it until all traces of color to the flame have disappeared. Now again moisten the wire and touch it to Azurite (No. 39). Hold the wire in the flame again and you will note a very pretty bluish green color, which is characteristic of copper.

EXPERIMENT 225—Testing Brass for the Presence of Copper.

Hold a small piece of sheet brass (1" x ½") with a pair of pliers and heat it over the flame of the alcohol lamp until it becomes red hot. You will notice a beautiful green coloration in the flame, identical with that observed when heating a piece of copper wire. After heating the piece of brass for a few minutes, remove it from the flame and allow a few drops of cold water to fall upon the metal. A dark coating of black copper oxide will be formed, proving that brass contains copper.

SILVER

Like copper, silver occurs free in nature, often in the form of mixtures or alloys with other metals. Among the silver minerals, the sulfide, argentite, is the most important source of this metal. Silver is frequently associated with the sulfide minerals of copper and lead.

Silver is the best of all conductors of heat and electricity. It is not oxidized in the air, but as we have learned, it does combine with sulfur, forming silver sulfide. The principal commercial uses of silver are in silverware for the home, jewelry and for coinage. As we have learned, the silver coins of our country contain 90% silver. Sterling silver contains 92.5% silver. Since pure silver is soft, it is alloyed with small amounts of copper to harden it.

Among the more important compounds of silver are silver nitrate and the silver halides, including silver chloride, bromide and iodide, which are used in the photographic industry because of their sensitiveness to light.

EXPERIMENT 226—Oxidized Silver.

"Oxidized" silver is silver which has been given a rich, dark tone by the formation of a layer of silver sulfide on its surface. A wide variety of techniques is used for doing this. You can do it very simply by placing 1 measure of Sodium Thiosulfate (No. 16) on a bright silver coin and heating it gently in your alcohol lamp flame. Hold the coin by one edge with your test tube holder. First the solid melts and then steam comes from it (its water of hydration). Slowly the surface of the silver darkens as sulfur from the thiosulfate combines with it. Rinse the coin when it has become as dark as you want it.

EXPERIMENT 227—How to Clean Silverware with Sodium Thiosulfate.

It may seem strange that the same compound with which you darkened silver in the preceding experiment can also be used to clean it. In order to darken silver, you heated the sodium thiosulfate. Now using sodium thiosulfate in solution, you will see that it dissolves the silver sulfide. Prepare a solution of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. Moisten a clean soft cloth with this solution and rub a piece of silverware which is tarnished with black silver sulfide. You will find that the silver sulfide comes off very easily.

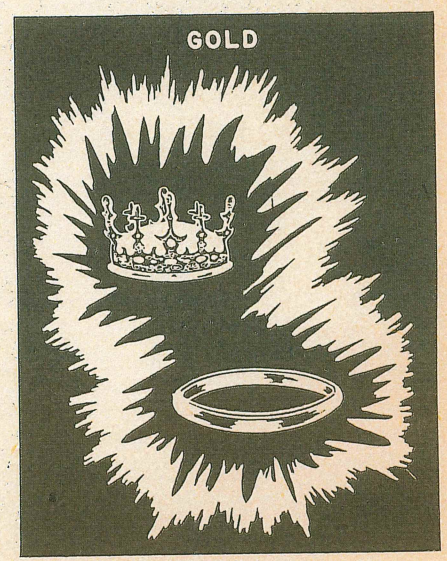
EXPERIMENT 228—The Electrolytic Cleaning of Silverware.

Obtain a piece of silver which is stained dark with silver sulfide. (If necessary expose the silver to some hydrogen sulfide gas.) Place the silver in an aluminum pan, cover with water and add common salt at the rate of one teaspoonful to each quart of water. Heat the pan on the stove until the water boils and the tarnish disappears from the silver. Only a minute or two of boiling should be necessary.

The aluminum is corroded while the silver is cleaned, so it is best not to leave the silver in the pan longer than necessary. If an old leaky aluminum pan is available, this may be used by setting it in a large enameled or agateware pan.

GOLD

From the beginnings of history, gold has been held as the symbol of wealth and as the king of metals. Its beautiful color and luster, its freedom from corrosion, the ease with which it is worked and shaped and finally, its rarity have caused it to be most highly prized. In more modern times gold has become the security for the monetary systems of the greater part of the world.



As might be expected, from its chemical inactivity, gold is generally found in the free state in nature, although often somewhat contaminated with other metals, especially silver. It occurs as veins in quartz rock and as this rock is disintegrated and washed down by streams, the gold accumulates in the stream bed as chunks called nuggets, and as flaky particles. Nuggets of gold weighing 100 lbs. or more have been discovered, but much more numerous are the very small fragments, ranging down to the size of dust specks.

Gold is the most malleable of all metals. With care, it can be hammered down to a thickness of three or four millionths of an inch. It is not attacked by any single acid, but a mixture of nitric and hydrochloric acids, known as aqua regia, dissolves gold, converting it into the chloride, known as auric chloride, $AuCl_3$.

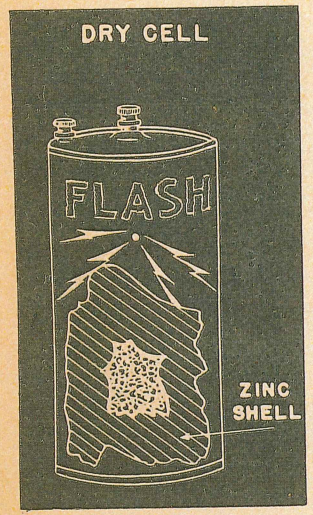
The principal uses for gold are coinage, jewelry, plating for ornamental work and in dental work. As pure gold is quite soft, it is usually alloyed with silver and copper to make it harder. Pure gold is 24 carats fine. Thus 18 carat gold would be 18/24 pure or 75% gold.

ZINC, CADMIUM AND MERCURY

Zinc, cadmium, and mercury constitute sub-group b of Group II of the periodic system, being thus cousins of the alkaline earth metals, to which they bear about the same relation as copper, silver, and gold bear to the alkali metals. These three elements all show a valence of two, but mercury sometimes has in addition a valence of one. Since zinc and cadmium resemble each other much more than mercury, we may discuss them together.

ZINC AND CADMIUM

The most common zinc minerals are the sulfide, ZnS, called zinc blende or sphalerite; the carbonate, $ZnCO_3$, called smithsonite; and a complex combination of the oxides of zinc, iron and manganese, known as franklinite. Cadmium is found in many zinc ores in small quantities.



In obtaining metallic zinc, the ores are roasted to convert them into zinc oxide. The oxide is heated with powdered coal, thus being reduced to metallic zinc, which comes off in the form of a vapor and is condensed in a receiver.

Zinc is a silvery white metal, but soon becomes bluish gray, due to the formation of a film of basic carbonate, formed by the action of oxygen and carbon dioxide in the air. At high temperatures, zinc burns in the air to form zinc oxide. It displaces hydrogen easily from acids with the formation of the corresponding salts of zinc.

Cadmium is very similar in physical and chemical properties to zinc.

The principal use of zinc is as a coating on iron and steel (galvanizing). This coating protects the iron from corrosion. Zinc is also used quite widely in various alloys, especially with copper. Metal for diecasting consists principally of zinc. Zinc is used in dry cell batteries, the shell or container of each cell often being made of zinc and serving as the negative electrode. Large amounts of zinc oxide are used as pigment for white paint.

Cadmium is extensively used as a protective coating on iron and steel. It is equal in protective power to zinc and it makes a better looking finish. Cadmium is also alloyed with other metals to make bearings and it is used in making the low melting alloys used in automatic sprinkler heads.

EXPERIMENT 229—Dissolving Zinc in Vinegar.

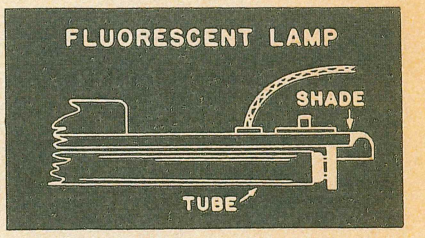
Fill a test tube ¼ full of ordinary vinegar and drop in a piece of zinc metal about the size of the end of your measure. Warm the vinegar gently to speed up the reaction. The bubbles which rise from the zinc are hydrogen gas. Set the test tube aside for an hour or two to give the acetic acid of the vinegar plenty of time to dissolve the zinc.

Meanwhile dissolve in another test tube ¼ full of water 1 measure of Sodium Carbonate (No. 4), and in a third test tube also ¼ full of water, dissolve 1 measure of Sodium Ferrocyanide (No. 6). After sufficient zinc has reacted with the vinegar, pour a few drops of the zinc solution into each of the other two test tubes. In the sodium carbonate solution you will find that the solution containing zinc acetate gives a white precipitate of basic zinc carbonate. In the sodium ferrocyanide solution a precipitate of white zinc ferrocyanide is formed. The fact that these precipitates are formed shows that zinc went into solution in the vinegar. This experiment demonstrates why galvanized utensils are not suitable for cooking acid foods.

MERCURY

Mercury is the only metallic element which is liquid at ordinary temperatures, although gallium is a liquid just above the usual room temperature. Mercury is heavier than lead, having a density of 13.6 grams per cubic centimeter. It dissolves many metals with the formation of alloys called amalgams. Some of these are solid and some are liquid at ordinary temperatures.

Mercury is sometimes found in the free state in nature often amalgamated with silver or gold. The most important source of mercury, however, is the mineral cinnabar, which is mercuric sulfide. This ore is roasted in the presence of air so as to burn out the sulfur and the metallic mercury distills off in the form of vapor and is condensed in a receiver.



The unique properties of metallic mercury make it useful for a great variety of purposes. Its application in thermometers and barometers is familiar. Amalgams of silver, gold and tin are used in dental fillings. When electricity is passed through mercury vapor at low pressure, light is emitted, which is very rich in ultra violet radiation. Thus mercury vapor lamps in a quartz tube, which is transparent to ultra violet light, are an important source of ultra violet radiation in chemical and therapeutic operations. Certain compounds are capable of absorbing the ultra violet radiations and re-emitting this light energy in visible form. Thus, if we coat the inside of a tube with such substances, called phosphors, fill the tube with mercury vapor and pass a current through it, we have the so-called fluorescent lamp, which is widely used in modern lighting.

Soluble mercury compounds are highly poisonous, for example, mercuric chloride, often called bichloride of mercury. Mercurochrome is a complex organic mercury compound which is used as an antiseptic on cuts and wounds.

EXPERIMENT 230—Mercury from Mercurochrome.

Fill a test tube ⅛ full of Mercurochrome solution and add an equal amount of water. Drop in 2 measures of Sodium Bisulfate (No. 7) and heat to boiling. Then add ½ measure of Bleaching Powder (No. 38) and agitate the test tube to mix in the bleaching powder. Continue to heat gently, being very careful not to let the slushy mixture spurt out. Soon the color begins to fade and after several minutes,

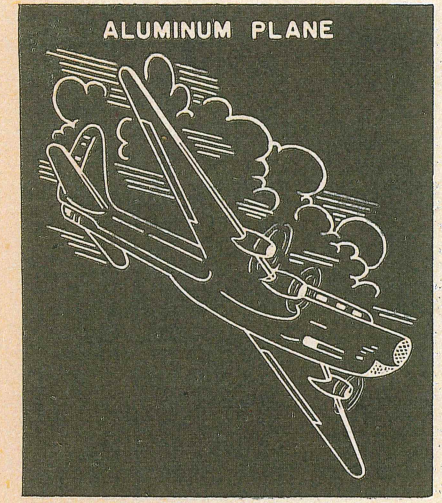
it should be almost entirely gone. If necessary, add one or more additional small portions of bleaching powder, but avoid putting in a large excess. Mercurochrome is a complex organic mercury compound and this treatment decomposes it and brings the mercury into the form of ions. When most of the colored material has been decomposed, pour the solution into a tumbler and immerse in it a clean, bright penny. The penny may be cleaned with steel wool or by boiling it in vinegar. After a few moments, the penny becomes silvery due to the mercury which deposits upon it, forming an amalgam with the copper. Rinse the penny and rub it with a soft cloth to make it shiny.

ALUMINUM

Aluminum is the most important metal in Group IIIa of the Periodic Table. The other metals in this group and also the metals in Group IIIb are all rarer elements which are of little importance commercially. Aluminum is trivalent, meaning that it has a valence of three.

Aluminum is the third most plentiful of the elements in the crust of the earth, the first being oxygen and the second, silicon. Aluminum is contained in many of our common mineral substances, including clay, feldspar, mica, garnet, corundum and bauxite. This latter, bauxite, is the chief commercial ore of aluminum and is a form of aluminum oxide.

In spite of the abundance of aluminum, the great chemical activity of the free metal makes its extraction in the pure state quite difficult. The first practical method of obtaining aluminum for its ores was discovered in 1886 by Charles M. Hall, a student at Oberlin College. The purified oxide of aluminum is melted in cryolite (a sodium aluminum fluoride) and an electric current passed through the solution. The melting point of aluminum oxide is too high to permit melting it directly for electrolysis, but it readily dissolves in melted cryolite and the aluminum ions in this solution are set free as aluminum metal in the cathode of the electric cell.



The most important use for aluminum is in the metallic state and as you know, many cooking utensils and articles about the home are made of aluminum. Industrially, aluminum is used in the construction of airplanes, trains, buses, truck bodies, or wherever light weight, strength and permanence are of importance.

Among the important compounds of aluminum is aluminum oxide from which the metal is made and which is also used in the manufacture of abrasives, either in the natural form known as corundum, or in an artificial form known as alundum. Aluminum oxide is also used as a refractory in lining high temperature crucibles and in making firebrick for industrial furnaces. Rubies and sapphires are composed of aluminum oxide, colored by traces of the oxides of other metals. Aluminum sulfate is used quite widely in clarifying water. Aluminum also forms a class of compounds known as alums, which are really double sulfates. Potassium alum, for example, is $K_2SO_4.Al_2(SO_4)_3.24H_2O$. Sodium alum is $Na_2SO_4.Al_2(SO_4)_3.24H_2O$. And ammonium alum is $(NH_4)_2SO_4.Al_2(SO_4)_3.24H_2O$.

EXPERIMENT 231—Dyeing Aluminum.

Recently it has been found possible to color aluminum with dyes by first forming a suitable coat of aluminum oxide on the surface.

Dissolve ½ measure of Sodium Carbonate (No. 4) in a test tube ¼ full of water and add 10 or 12 drops of ordinary hydrogen peroxide. Dip into this solution a small piece of aluminum (cut from an old utensil) and boil several minutes. This will produce a thick coating of aluminum oxide on the metal.

In another test tube place 5 or 6 drops of blue ink and fill the tube ¼ full of water. Now remove the aluminum strip from the oxidizing bath, wash it in running water and immerse it in the dye solution. After 5 minutes take it out, wash again in running water and let it dry.

EXPERIMENT 232—Dyeing Aluminum Red.

Repeat the preceding experiment, but instead of the blue ink, use ½ test tube full of red ink. You will find that the strip of aluminum metal takes up red coloring matter from the ink.

EXPERIMENT 233—Dyeing Aluminum Green.

By using green ink you will find that the aluminum oxide takes up green dye where the oxide is sufficiently thick.

TIN, LEAD AND OTHER METALS OF GROUP IV

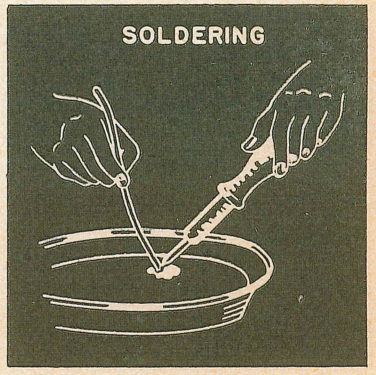
Tin and lead, together with germanium constitute Group IVb of the Periodic Table, while titanium, zirconium, hafnium and thorium constitute Group IVa. The important metals in Group IV are, of course, tin and lead. Both of these metals may possess a valence of either 2 or 4. Tin and lead are typically metallic from the standpoint of physical properties.

TIN

Tin is a relatively scarce metal and unfortunately, the United States has no important deposits of tin ores. The only important ore of tin is cassiterite, or tinstone, which is stannic oxide, SnO_2. The most important sources of this ore are Malaya, the Dutch East Indies and Bolivia. In producing metallic tin, the ore is roasted and then smelted or reduced with carbon in the form of coal or coke,

$$SnO_2 + 2\,C \rightarrow Sn + 2\,CO$$

Tin is a relatively inactive metal and hence is rather easily obtained from its oxide. Conversely, its oxide is not very readily formed, and so tin is fairly resistant to corrosion.



The greatest use of metallic tin is in coating steel to protect it from corrosion. Our so-called "tin" cans contain only 1% or less of tin deposited in an extremely thin film over the surface of the steel. This coating was formerly done by passing the steel through a bath of melted tin. In order to save tin, however, most of the coating is now done by electroplating. Tin is also of importance in many alloys. Solder contains various

proportions of tin and lead. Type metal consists of lead, antimony and tin. Babbitt metal used in bearings to reduce friction contains copper, antimony and tin.

LEAD

Lead is a soft heavy metal having a specific gravity of 11.3. The principal source of lead is lead sulfide, PbS, the mineral galena. The concentrated lead ore is roasted to burn out part of the sulfur, leaving a mixture of lead oxide and lead sulfide. This is mixed with coke and limestone and charged into a blast furnace. Lead sulfide and lead oxide react, setting free the lead. The coke also acts as a reducing agent. The purpose of the limestone is to combine with any silica which may be present in the ore.

Due to its resistance to corrosion and also to the plentiful supply and low price, lead has a large number of uses. The acid type of storage battery uses lead plates. Telephone cables are sheathed in lead to protect them. Much equipment in chemical plants is constructed with lead or lined with it since it is resistant to chemical attack. Lead shot and bullets are familiar. Lead pipe is used extensively for drains. Lead is used in a number of alloys, as previously mentioned.

Some of the compounds of lead are of great importance industrially. Lead monoxide, PbO, in which the valence of lead is 2, is commonly called litharge. Litharge is used in paints as a drying agent and it also forms a waterproof cement when mixed with glycerin. Lead dioxide, PbO_2, in which lead has a valence of 4, is the active substance on the positive plate of a charged storage battery.

White lead is one of our most valuable white pigments for paint. It is a basic carbonate, that is, a combination of lead hydroxide and lead carbonate. Red lead, Pb_3O_4, is commonly used in paints for structural steel. Lead arsenate is an important insecticide, and tetraethyl lead is the anti-knock fluid added to ordinary gasoline to make ethyl gasoline.

ARSENIC, ANTIMONY AND BISMUTH

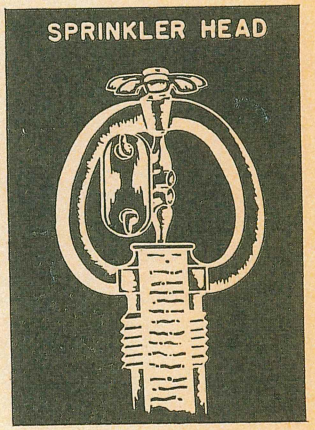
Arsenic, antimony and bismuth are classified in Group V, sub-division-b of the Periodic Table, while vanadium, columbium and tatalum are classified in Group Va. The most important metals in these groups are arsenic, antimony and bismuth. These metals each possess a valence of either 3 or 5 and in addition, antimony may in rare cases be 4 valent. These three elements, although they look like metals, possess many of the properties of non-metals. The oxide of arsenic, for example, when dissolved in water, forms an acid instead of a base. The oxides of antimony and bismuth, however, when dissolved in water, form weak bases. The property of arsenic mentioned above accounts for the fact that its oxide can combine with bases, forming arsenates.

ARSENIC

Free arsenic has few uses, the chief being as an alloying constituent with lead in making shot where it tends to harden the lead. By far the greatest use for arsenic compounds is in the preparation of poisons for rodents and insects. The largest amounts of arsenic probably go into the manufacture of lead arsenate and for calcium arsenate for use in insect sprays.

ANTIMONY

The greater metallic character of antimony is reflected in its use in alloys to which it imparts the very valuable property of expanding slightly upon solidifying. This is especially important in the casting of type for printing since it causes the solidified metal to fill the mold completely, giving a sharp impression. Like arsenic, antimony has the effect of hardening lead and also of increasing its resistance to corrosion. Hence, lead storage battery plates and lead sheaths for telephone cables contain some antimony. The sulfides of antimony, when prepared by precipitation, are orange in color and are used as a pigment in coloring rubber goods.



BISMUTH

The most important and interesting use of bismuth is in low melting point alloys. The melting point of solid bismuth is 271° C., but when we dissolve in the bismuth 25% of lead and 12½% each of tin and cadmium, the melting point of the alloy is lowered to 70° C., much below the boiling point of water and far below the melting point of any of the added metals. This alloy is used in the heads for automatic sprinkler systems, arranged so that when the alloy is melted by excessive heat, a valve is released and the water is turned on.

CHROMIUM, TUNGSTEN, URANIUM AND MANGANESE

In this section we shall discuss several less abundant, but nevertheless, important metals. Among them are representatives of several groups of the Periodic Table and the reactions of most of them are rather complicated since they may exist in several valence states.

CHROMIUM

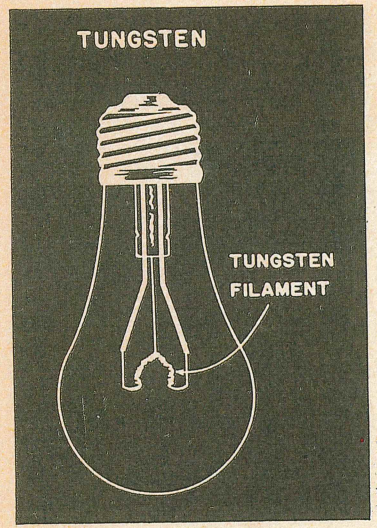
Chromium is classified in the Sixth Group, sub-division a, of the Periodic Table and its valence may be either 2, 3, or 6.

The most important use for chromium is in alloys with steel to prevent corrosion. Typical stainless steel contains about 18% chromium along with 8% nickel. Chromium is also electroplated on various metal parts. It produces a bright surface of great hardness and strength, which is highly resistant to corrosion.

The most important compounds of chromium are the chromates and dichromates. These compounds are the basis of certain yellow and orange pigments used in paints and they are also used in tanning leather, as chrome-tanned hides make the most waterproof leather.

TUNGSTEN

This element, together with molybdenum, is classified in the same group in the Periodic Table as chromium. Both of these elements are of great importance in alloy steels. Tungsten has the highest melting point of any



metal — 3370° C. As obtained by reduction of its oxide, tungsten is a fine powder. In order to form this metal into wire, it is heated under great pressure which causes it to sinter together sufficiently so that it can be worked into thinner and thinner rods and finally drawn into wire. This wire is used in incandescent lamp filaments and in radio and x-ray tubes.

URANIUM

Uranium is classified in the Periodic Table among the Rare Earths. It has the highest atomic number of any element found in nature.

The principal use for uranium was formerly in uranium glass, to which it imparts a yellowish color. Due to the remarkable reactions which its nucleus undergoes, uranium is now in large demand as a source of atomic energy.

MANGANESE

Manganese is classified in Group VII, sub-division a, of the Periodic Table. It may exist in five valence states, namely, 2, 3, 4, 6, or 7 valent. This results in the formation of five well defined sets of compounds arising from the five oxides of manganese. The lower oxides of manganese, MnO and Mn_2O_3, are basic, while the higher oxides, MnO_3 and Mn_2O_7, when dissolved in water are acid. The intermediate oxide, MnO_2, called manganese dioxide, is practically neutral, but shows very weakly acidic properties.

The most important compound of manganese is manganese dioxide, in which state it occurs in nature as the mineral, pyrolusite. Manganese is used as an alloy with steel, as it has the property of making steel very hard and tough.

An important use of manganese dioxide is in the production of ordinary dry cells, where it serves as a depolarizer to prevent, by its oxidizing action, the accumulation of hydrogen on the carbon pole of the cell.

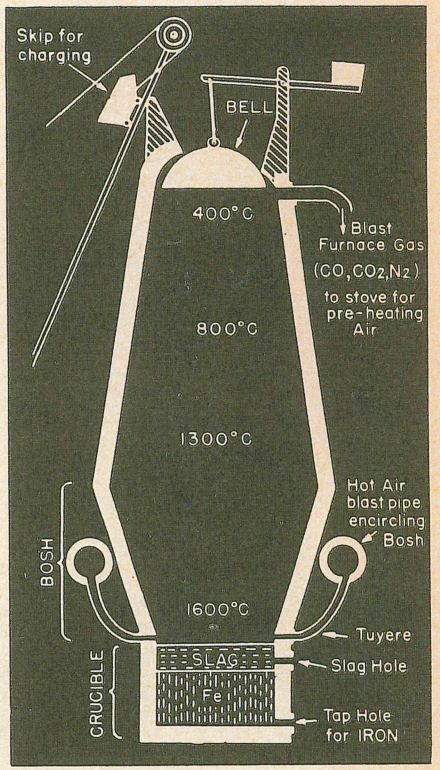
IRON, COBALT, NICKEL AND THE PLATINUM METALS

The metallic elements considered in this section are known as Transition Elements and appear under this heading in the Periodic Table. Iron, the most important metal in this group, commonly exhibits a valence of 2 or 3, but may in rare cases have a valence of 4 or 6. Cobalt likewise is commonly either 2 or 3 valent but may exhibit a valence of 4. Nickel is commonly 2 valent but in certain cases may show a valence of 3 or 4. Platinum exhibits a valence of 2 or 4 and may in rare cases have a valence of 3 or 6. Iron, cobalt and nickel resemble each other very closely.

IRON

Iron is the fourth most abundant element in the crust of the earth, following oxygen, silicon, and aluminum. It is thought that the core of the earth may be composed almost entirely of iron and nickel.

The most important ores of iron are all oxides and the most abundant of these is the mineral hematite, Fe_2O_3. Iron is obtained from its ores by a process of reduction carried out in a blast furnace. The reducing agent is carbon or rather the carbon monoxide obtained from it. The blast furnace is a great steel tower, about 100 feet high and 20 feet across, lined with fire-brick. Charges of iron ore, coke and limestone are added through the top, while preheated air is blown in at the bottom. The reaction is one of reduction in which the iron oxide is reduced to metallic iron and the carbon, represented by the coke, is oxidized by the oxygen contained in the iron ore plus the oxygen in the air. Limestone is used to react with the silica and alumina which are impurities in the ore and forms a slag consisting of calcium silicate and calcium aluminate. The molten iron metal is drawn off at the bottom of the furnace.



Cast iron results when the impure product of the blast furnace is melted with scrap iron. This cast iron is relatively brittle but because of its cheapness, it is used for heavy machine bases, automobile cylinder blocks, steam and hot water radiators and many similar objects. Cast iron contains about 2% of carbon and part of this carbon is in the form of graphite interspersed with the grains of iron. The rest of the carbon is combined with the iron in the form of iron carbide.

EXPERIMENT 234—To Show the Presence of Free Carbon in Cast Iron.

Place ½ measure of powdered iron in a test tube. Add 6 measures of Sodium Bisulfate (No. 7) and fill the test tube ⅓ full of water. Warm gently until the sodium bisulfate is completely dissolved.

You will notice that the iron is being dissolved by the sodium bisulfate solution and that hydrogen gas is driven off. Mixed with this hydrogen is a small amount of hydrocarbon gases having an unpleasant odor. These gases are formed from the small amount of combined carbon in the cast iron. The free carbon or graphite, however, is not affected by the sodium bisulfate and remains as a black powder after all the iron has been dissolved.

STEEL

The term steel covers a number of different types of iron alloys and it is difficult to give an accurate definition. One of the reasons why iron is so useful is that its properties may be modified through a wide range by the use of small quantities of alloying elements. Chief among the elements which affect the characteristics of steel is carbon. When the carbon content is low (less then ½%), the steel will be relatively soft. Higher carbon content makes it possible to harden steel through heat treatment.

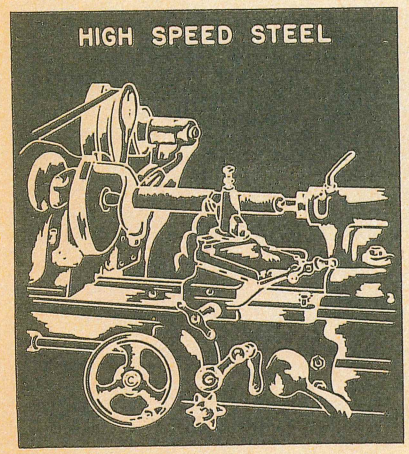
EXPERIMENT 235—Heat Treatment of Steel.

Heat a discarded safety razor blade to a bright red heat in the flame of an alcohol lamp or of a gas stove and gradually withdraw it from the flame so that it cools slowly. Then try bending it. You will find that it has lost most of its springiness. Suitable tests would show also that it is no longer very hard. You have "drawn the temper" of the steel.

Now heat the blade again, but this time cool it very suddenly by quenching in a vessel of cold water. After this treatment, you will find that the blade is extremely brittle and can be bent scarcely at all without breaking. It is, nevertheless, very hard. In making razor blades and similar articles, they are first hardened in this way by quenching and then heated to a carefully controlled temperature and allowed to cool slowly so as to soften them just enough to remove most of the brittleness. This also makes them less hard, of course, but some of the desirable hardness must be sacrificed for the sake of reducing the brittleness. As you did the experiment, a considerable amount of oxide scale was probably formed. To avoid this in commercial operations, the heating is frequently done in the presence of non-oxidizing gases.

Now try hardening, by the same process, a thin iron nail and see whether you can harden it enough so that it breaks when you bend it. You will probably not succeed, since nails are made of low-carbon steel, and unless at least 0.5% carbon is present, the steel can not be hardened by heat treatment.

Manganese will make steel tougher and harder. Nickel and chromium also give steel greater toughness and hardness. Relatively large amounts of nickel or chromium or both (up to about 20%) will produce stainless steel, which is not subject to oxidation or corrosion. A fraction of a per cent of vanadium increases the tensile strength of steel and is often used in spring steel. Tungsten makes steel very hard, even when it is red hot, so tungsten is used in so-called high-speed steels for making tools for machining metals. Usually chromium, vanadium and cobalt are used in connection with tungsten. Other elements sometimes used in alloy steels are copper, silver, lead, titanium, columbium, cerium, uranium, aluminum, tantalum, and no one can foretell what other elements may be found to give beneficial properties to steel, the backbone of our civilization.



You are familiar with several of the compounds of iron which have been used in previous experiments. The compounds of this metal, however, do not approach in importance the metal itself.

COBALT AND NICKEL

These metallic elements occur in the crust of the earth only to the extent of a small fraction of 1%, however, as mentioned, large quantities of nickel may be present in the center of the earth. As we have learned, both nickel and cobalt are of importance in steel where they impart special characteristics to the metal. Cobalt compounds are used to color glass blue and to impart a beautiful color to pottery glazes.

Nickel is used extensively in electroplating as metallic nickel does not oxidize or tarnish under ordinary conditions and it forms a hard bright durable finish on other metals. The principal use for nickel compounds is in plating baths where the nickel ions move toward the negative electrode and deposit on the work to be plated.

PLATINUM

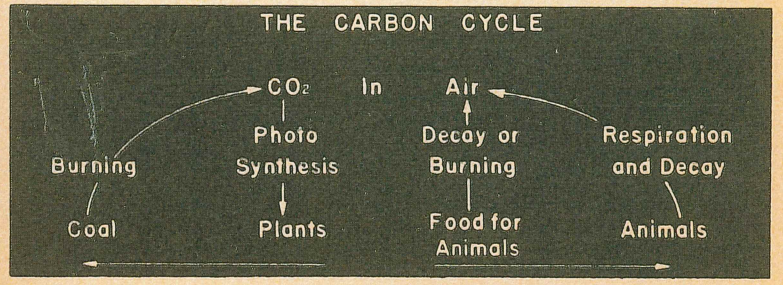
Platinum is the principal metal in a group of six metallic elements, all of which have quite similar physical and chemical properties. These elements are ruthenium, rhodium, palladium, osmium, iridium and platinum.

Platinum is used in jewelry, in dentistry and in chemical industries and laboratories. Platinum crucibles and evaporating dishes are indespensable because of their resistance to chemical attack. In finely divided form, platinum is a valuable catalyst in the contact sulfuric acid process and in the oxidation of ammonia to nitric acid.

Life On Our Earth

We live in a chemical world and we ourselves, as well as plants, animals and every living thing, depend on chemistry. We have learned that as we breathe, the oxygen of the air unites with carbon in our bodies forming carbon dioxide, which we exhale in our breath. Plants, on the other hand, absorb carbon dioxide from the air and give off oxygen, thus, both plant and animal life exist by chemical reactions.

The fact that all animal life, including man, adds to the supply of carbon dioxide in the air and that all combustion of coal, oil, gasoline and other fuels also adds to the supply of carbon dioxide, might lead us to believe that eventually the oxygen in our atmosphere would be all used up and that animal life could no longer exist. Counterbalancing this enormous use of



oxygen and production of carbon dioxide is the absorption of carbon dioxide by plants which retain carbon and release oxygen to the air. Thus, animal and plant life are mutually dependent upon each other to maintain a proper balance in our chemical world.

The leaves of green plants have the power of combining carbon dioxide and water to form sugars, starches, and cellulose. This reaction cannot take place unless energy in the form of light is supplied to the leaf, and hence the process by which plants use carbon dioxide and water is called "photosynthesis." Without this reaction animal life as we know it would be impossible. It was by means of photosynthesis that the energy of the sun's rays millions of years ago was stored in the form of coal. The carbon in coal was originally obtained from the carbon dioxide of the air by the energy of sunlight shining on green leaves.

Thus plants convert the carbon dioxide of the air into substances which can be used for food or fuel, and then the food is oxidized in the bodies of animals and the fuel is burned in a furnace, converting the carbon back again to carbon dioxide. The diagram shows the more important stages in this cycle chain of interconversions.

We have already learned in a previous experiment that leaves give off oxygen in the sunlight. Now let us demonstrate that carbon dioxide is present in the breath we exhale.

EXPERIMENT 236—Showing That There Is Carbon Dioxide in Our Breath.

Fill an ordinary tumbler or drinking glass ¼ full of water and add 1 or 2 drops of Phenolphthalein Solution (No. 22). In a clean test tube dissolve 1 measure of Sodium Carbonate (No. 4) in half a test tube of water. After the sodium carbonate is all dissolved, add 2 or 3 drops of the solution in the test tube to the glass containing the phenolphthalein and water. Add just enough of the sodium carbonate solution to turn the solution in the glass a pink color. It will take only a few drops.

Now, using your glass tube, blow into this pink solution in the glass. Inhale through your nose and exhale through the glass tube so that your breath bubbles through the pink solution. After you have done this for a few minutes, you will notice that the pink solution is gradually becoming lighter and after a while it will be entirely colorless. What has happened?

Since sodium carbonate is the salt of a very weak acid, carbonic acid, its solution is alkaline and that is why it turned the phenolphthalein solution pink. We also know that carbon dioxide is an oxide of a non-metal (carbon) and that when this non-metallic oxide is dissolved in water, it forms an acid (carbonic acid). Thus, as we blow our breath through the sodium carbonate solution, the carbon dioxide gradually dissolves in the water, forming carbonic acid. This acid converts the alkaline sodium carbonate into sodium bicarbonate, which contains a sufficient excess of the weak acid to make its solution nearly neutral. Thus, the phenolphthalein becomes colorless.

EXPERIMENT 237—Another Way of Proving the Presence of Carbon Dioxide in Our Breath.

Prepare some lime water by adding 2 measures of Calcium Oxide (No. 20) to a glass half full of water and stirring for a minute or two. Let the undissolved solid particles settle to the bottom and pour the clear solution into another glass.

Using a piece of glass tubing, blow through the lime water. You will find that as your breath bubbles through the liquid, it causes it to turn white and turbid, thus showing the presence of carbon dioxide.

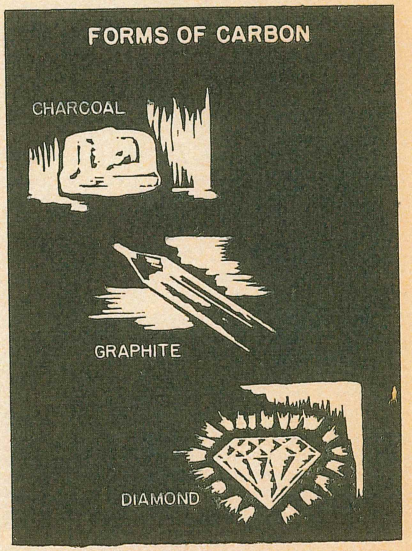
The white precipitate is calcium carbonate, which is soluble in most acids. Stir in a little vinegar and watch the turbidity disappear.

CARBON

Carbon is the first element in Group IVa of the Periodic Table. It may exist in the valence state of 2 or 4, but almost always exhibits a valence of 4. Carbon is classified as a non-metallic element because its oxide unites with water forming an acid.

Carbon, although it makes up less than a fifth of 1% of the earth's crust, is one of the most important elements because it is a constituent of so many valuable compounds and because it enters into numerous important reactions. It occurs in the free state as diamonds and as graphite. Many carbonates occur as rock-forming minerals, especially limestone and dolomite. Coal and petroleum are essentially mixtures of carbon compounds and all organic matter, plants and animals, consists in large part of carbon compounds. Furthermore, we find carbon in the form of carbon dioxide distributed through the atmosphere and dissolved in the water of the oceans.

Free carbon may exist in three allotropic forms. You will recall that sulfur and phosphorus exhibit similar properties. The allotropic forms of carbon are the diamond and graphite (both of which are crystalline in nature) and amorphous carbon, which may occur as coke, charcoal or lampblack.



EXPERIMENT 238—Forms of Carbon.

Examine a piece of graphite in a soft lead pencil. Notice its softness and smooth, greasy feel. It can be scratched even with your fingernail and rubs off onto paper, leaving a black smudge.

Now compare a piece of charcoal or coke with the graphite. It is harder and feels rough and harsh, although, it, too, will leave a mark on paper. Notice also that the charcoal is much less dense because of its many minute pores filled with air.

If you have a diamond ring, you can compare the most beautiful form of carbon with the other two. The transparent sparkle of the diamond is quite different from the dirty blackness of graphite and charcoal. Draw a corner of the diamond over a piece of glass. Notice how readily the glass is scratched. Graphite is one of the softest minerals. The diamond is the hardest of all, yet both are forms of pure carbon.

COMPOUNDS OF CARBON

Carbon forms many more compounds than any other element. Hundreds of thousands of compounds of carbon are already known and new ones are being discovered continually. It is no wonder then that the chemistry of the compounds of carbon has been set aside as a separate and very important branch of chemistry—organic chemistry. This tremendous variety of com-

pounds formed from carbon is possible because carbon atoms, unlike those of most elements, are able to unite with one another, building up long chains which may then join into rings or have intricate branches attached to them, with other elements joined on at various places. These complicated carbon compounds studied by the organic chemist are of the very greatest importance. Some of them are fundamental to all living things. Others are dyes, drugs and explosives. Foods, textiles, wood, paper, rubber, plastics, coal and oil are all composed of organic compounds. We will first consider some of the simpler compounds of carbon, which are studied in inorganic chemistry.

CARBON MONOXIDE

When carbon burns in a limited amount of air, each atom of carbon unites with one of oxygen to form carbon monoxide:

$$\underset{\text{(carbon)}}{2\,C} + \underset{\text{(oxygen)}}{O_2} \rightarrow \underset{\text{(carbon monoxide)}}{2\,CO}$$

This compound is a gas, and it is extremely poisonous when inhaled. Since it is present in small amounts in the exhaust gases from automobile engines, it is well to be careful about breathing such gases, especially in closed garages. In particular, never run an automobile engine without opening the garage doors. Many people have lost their lives by disregarding this simple rule.

CARBON DIOXIDE

When carbon burns with an adequate supply of oxygen, each atom combines with two atoms of oxygen, and carbon dioxide results:

$$C + O_2 \rightarrow CO_2$$

It is also formed when carbon monoxide burns, taking up another atom of oxygen. A more convenient way of making it, however, is to treat a carbonate with an acid.

EXPERIMENT 239—Making Carbon Dioxide.

To a test tube ⅔ full of water, add ½ measure of Calcium Oxide (No. 20) and shake it vigorously for a few minutes. Let the undissolved solid settle to the bottom and pour off the clear liquid into a third test tube. This is lime water, a solution of calcium hydroxide.

Place 3 measures of baking soda in a test tube. Add ¼ test tube of vinegar and quickly fit the gas delivery tube to the mouth of the test tube, insert the long end of the delivery tube into the lime water, and let the carbon dioxide gas which is set free from the baking soda pass through the tube into the lime water. When the evolution of gas slackens, remove the stopper for an instant and drop in another measure of baking soda.

Soon you will see that the clear lime water is becoming turbid or cloudy, because the carbon dioxide is reacting with it to form insoluble particles of calcium carbonate. This is a positive test for the presence of carbon dioxide.

EXPERIMENT 240—Carbon Dioxide Does Not Support Combustion.

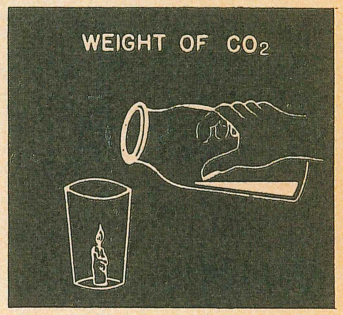
Dissolve 6 measures of baking soda in a test tube half full of water. Obtain 2 tumblers or jelly glasses. Make a cardboard or heavy paper cover to fit over the top of one of the glasses and in this cover, make a hole about ¼" in diameter.

Have your gas delivery tube and stopper handy. Now add ½ measure of Sodium Bisulfate (No. 7) to the solution of baking soda in the test tube. Quickly insert the gas delivery tube and stopper and place the end of the gas delivery tube through the hole in the piece of cardboard covering one of the tumblers. Thus, the carbon dioxide which is being generated in the test tube is directed into the tumbler. As soon as the effervescence or action in the test tube stops, add another half measure of sodium bisulfate. Again insert the gas delivery tube and stopper and continue to collect the gas in the tumbler. Repeat this until no further effervescence occurs when you add sodium bisulfate. Now strike a match and holding it carefully, lower it into the empty tumbler. If the match is immediately removed, it will not go out, but continues to burn. Now strike another match and holding it the same way, remove the cardboard cover and lower it into the tumbler containing carbon dioxide gas. The match will go out immediately, showing that carbon dioxide does not support combustion.

EXPERIMENT 241—Weight of Carbon Dioxide Gas.

Carbon dioxide is much heavier, that is, much denser, than air. Its high density makes it possible to pour it from one vessel to another almost like a liquid.

Obtain a quart milk bottle and in it place 2 teaspoonfuls of baking soda. Next set a short candle upright in the bottom of a tumbler or jelly glass, using a little melted wax to make it stick to the bottom of the tumbler. The wick should not come higher than half way up the tumbler. Pour about half a glass of vinegar into the milk bottle and while the carbon dioxide gas is being formed and pushing the air out of the bottle, light the candle. Hold the milk bottle over the glass and pour carbon dioxide out of it onto the candle just as if you were pouring a liquid, but be careful not to allow any of the vinegar to pour out. When you have filled the glass up to the level of the flame with carbon dioxide, the candle will go out, for the oxygen supply will be cut off by the heavy CO_2 gas.



Carbon dioxide, because of its relatively high density, tends to collect in abandoned mines and old wells, and so such places should always be tested by lowering a lighted candle into them before anyone descends. Sometimes carbon dioxide is formed by fermentation and accumulation in silos, and death by suffocation may result from entering and remaining in a closed silo.

We have found that carbon dioxide is a heavy gas that will not burn or support combustion. We know that fire will not burn when its supply of oxygen in the air is cut off. Therefore, we reason that carbon dioxide would be a good fire extinguisher. This is true and we find that carbon dioxide is used in many commercial fire extinguishers.

EXPERIMENT 242—Extinguishing a Flame with Carbon Dioxide.

Set a candle in a tumbler so that when lighted the top of the flame will be just below the edge of the tumbler. Now put a spoonful of baking soda in the tumbler and add vinegar, a few drops at a time, being careful to pour the vinegar in so that it will not get on the candle flame. In a few moments the carbon dioxide generated in the bottom of the tumbler will suffocate the candle and the flame will be extinguished.

Some hand fire extinguishers and many chemical fire engines make use of carbon dioxide. They contain a solution of sodium bicarbonate (baking soda) and sulfuric acid, held in separate containers, which are so arranged that when the extinguisher is turned upside down, the chemicals are mixed and carbon dioxide is generated. The pressure of this gas forces a mixture of water and carbon dioxide out of the nozzle in the form of a powerful stream which may be directed upon the fire.

EXPERIMENT 243—Making Candles Suffocate Each Other.

Obtain a quart milk bottle or fruit jar. Now fasten 3 candles to a wire, one above the other, so that when lowered into the bottle, one candle will be near the

bottom, the second about half way up and the third two or three inches below the top.

Light the candles and lower them into the bottle and cover the mouth with a piece of cardboard. As the candles burn, carbon dioxide is formed and since the gas is hot, it rises and collects at the top of the bottle, first extinguishing the upper candle and then those lower down.

EXPERIMENT 244—A Dry Fire Extinguisher.

Place a spoonful of baking soda on a sheet of tissue paper about 4 inches square and wrap the soda in the paper. Put this package on a tin plate and light the paper with a match. Notice that the flame goes out almost at once.

When heated, sodium bicarbonate or baking soda, gives off carbon dioxide just as it does when treated with an acid. The carbon dioxide set free by the heat of the flame promptly settles around the flame and extinguishes it.

EXPERIMENT 245—A Foam Fire Extinguisher.

Pour about half a test tube of thin liquid glue or mucilage into a glass and add 5 or 6 measures of sodium bicarbonate. Stir this mixture thoroughly so that it makes a thin, uniform paste. Now dissolve 4 measures of Sodium Bisulfate (No. 7) in a test tube ½ full of water. Shake the tube vigorously until the solid is all dissolved. Pour the sodium bisulfate solution into the glass containing the paste and a heavy, sticky foam will rapidly form. The bubbles in this foam are filled with carbon dioxide gas and the glue or mucilage makes it so stiff and sticky that it will form a coating on burning wood and smother the fire.

EXPERIMENT 246—Fireproofing Cloth.

Cloth and other inflammable substances may be fireproofed by treating them with chemicals which, when heated, give off vapors that smother the flame.

Dissolve by warming 12 measures of Ammonium Chloride (No. 9) in a test tube ⅓ full of water. Put a piece of cotton cloth, 2 or 3 inches square, in the bottom of the glass and pour the liquid in the test tube over it. Stir the cloth around until it is wet through, then let dry and try to light it with a match. You will find that it will burn while held in the flame, but just as soon as the match flame is removed, it will go out.

Because fabrics treated in this manner cannot be set on fire by sparks or overturned candles, a similar process of fireproofing is sometimes used for treating the curtains and scenery of theaters.

EXPERIMENT 247—Fireproofing Wood.

Sodium silicate solution (water glass) is sometimes used for fire proofing wood, since when it dries, it forms a hard, dense film over the wood through which oxygen cannot penetrate. It is not suitable for cloth or paper because it makes them stiff.

Hold a match by the head and dip about two-thirds of the stick into Sodium Silicate Solution (No. 19). Let the coating dry, and then light the match. There will be no danger of burning your fingers, for the flame will go out as soon as it reaches the film of sodium silicate.

SOLID CARBON DIOXIDE — DRY ICE

Carbon dioxide is a gas under ordinary conditions, but when cooled sufficiently, becomes solid. By compressing the gas, it is possible also to obtain liquid carbon dioxide, but the liquid cannot exist at the pressure of

the atmosphere, and when allowed to escape from a tank under pressure, it evaporates so rapidly that it absorbs enough heat to freeze some of it to the solid state. This is the usual method of making solid carbon dioxide.

Since when the solid absorbs heat it passes directly into the form of gas without first melting to a liquid, it is commonly called dry ice. Large quantities of it are now made for use as a refrigerant in the packing of ice cream and other foods as well as for refrigerator cars. The carbon dioxide used for making it is often obtained as a by-product, for example, from lime kilns or from fermentation vats.

To perform the experiments of this section you will need some dry ice. This may usually be obtained from a soda fountain or ice cream manufacturing plant, but since it evaporates rather rapidly, you had better not get it until you are ready to begin the experiment.

CAUTION: Dry ice can be handled safely with the bare hands so long as it is not squeezed. When pressed against the skin, however, heat is removed so rapidly from the area in contact with the dry ice, that a frostbite is produced which feels very much like a burn. Hence, it is best to handle this material with a pair of pliers or when wearing a leather glove.

EXPERIMENT 248—Testing Dry Ice for Carbon Dioxide.

Stir 2 measures of Calcium Oxide (No. 20) in half a glass of water for a minute or two and then let the undissolved solid settle out. Pour off the clear lime water into another glass, and drop into it a small piece of solid carbon dioxide.

At once the lime water becomes turbid, because of the calcium carbonate formed. This, you will remember, is the usual test for carbon dioxide. If enough carbon dioxide is present, the white precipitate of calcium carbonate may later disappear because an excess of carbon dioxide converts it into soluble calcium bicarbonate.

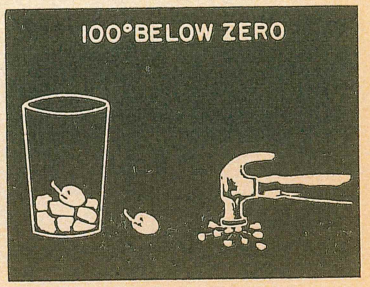
EXPERIMENT 249—How Cold Is Dry Ice?

In a previous experiment we found that although it was easy to freeze water in an ice and salt bath, a strong salt solution had too low a freezing point to solidify in such a cooling bath. We now have, however, a much more powerful cooling agent at our disposal.

Add about 10 measures of table salt to a test tube ½ full of water and shake the tube until the salt is dissolved. Set the test tube in a tumbler containing several pieces of dry ice upon which you have poured some ordinary denatured, or rubbing, alcohol. The alcohol makes better contact with the test tube than the dry ice can, and being itself cooled by the dry ice, it then removes heat rapidly from the test tube. You will soon find that the salt solution has solidified completely.

EXPERIMENT 250—100° below Zero.

Not only salt solution, but many other substances are frozen solid by dry ice. Crush some dry ice in the bottom of a tumbler, using a wooden stick, such as the handle of a hammer. Pour in some ordinary denatured alcohol and place in the resulting mixture various objects, such as a green leaf, a grape, a rubber band, a fragment of raw meat, and a small flower. After several minutes they will be thoroughly chilled. Remove them and examine them quickly before they warm up. You will find that the leaf and the flower have become so brittle that you can crush them into bits. The rubber is very stiff and perhaps even brittle. The grape and the meat are hard like pebbles, and if struck with a hammer while they are still cold, they will crumble into fragments. Of course you must not eat anything which has been soaked in denatured alcohol.



The temperature of the alcohol and dry ice mixture is about —75° C. (—103° F.). Cold though it is, dry ice is still hot compared to liquid air, the temperature of which is about —190° C. (—310° F.).

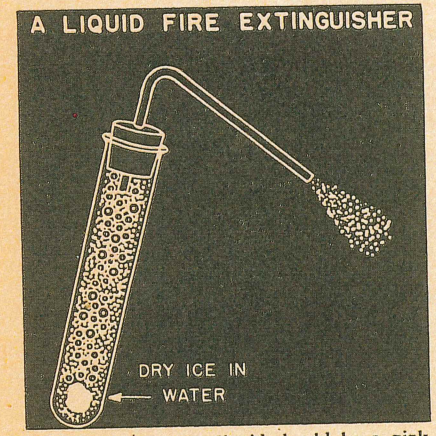
EXPERIMENT 251—Making Soap Bubbles Float on Carbon Dioxide.

Place a few pieces of dry ice in the bottom of a shallow dish, such as a soup plate. In another dish make some soap suds by stirring soap shavings in half a glass of warm water until dissolved. Using a soap bubble pipe, blow some bubbles and allow them to fall into the dish containing the dry ice. If a sufficient amount of the dry ice has evaporated, the dish will be filled with carbon dioxide gas, and since it is considerably more dense than air, the air-filled bubbles will float on it, just as a hollow rubber ball floats on water. Of course you must do this experiment in a room where there are no drafts, since otherwise the carbon dioxide would be blown out of the dish.

EXPERIMENT 252—Extinguishing a Candle with Dry Ice.

Place several pieces of dry ice in the bottom of a ½ gallon jar and then light your candle and place it upright on a tin pan on the table. After about a minute, the carbon dioxide gas formed by the evaporation of the dry ice will have filled the jar. Pick it up and pour the gas quickly over the candle flame, just as if you were pouring water. Being thus flooded with carbon dioxide, the candle can no longer get oxygen, and so the flame is extinguished.

EXPERIMENT 253—Making a Liquid Fire Extinguisher.



Fill a test tube ¾ full of water and drop into it several small lumps of dry ice. Quickly fit the test tube with your gas delivery tube and direct the end of the tube at a can or other convenient receptacle. A strong stream of water will be forced out of the tube by the pressure of the expanding gas. This is the same principle that is employed by the ordinary carbon dioxide fire extinguisher except that in the commercial variety the carbon dioxide is formed by the interaction of sodium bicarbonate and sulfuric acid as we have already learned.

EXPERIMENT 254—Making Carbonic Acid from Dry Ice.

Fill a tumbler half full of water, and add a very small quantity of Sodium Carbonate (No. 4) (⅛ measure should be enough). Now add 2 drops of Phenolphthalein Solution (No. 22). The liquid should be a pink color. Drop in several small pieces of dry ice, and as they evaporate and the gas dissolves in the water, the pink color of the phenolphthalein will fade away, showing that the originally basic solution has now become neutral and perhaps acid. Test it with blue Litmus Paper (No. 18) and you will find that the paper turns red, showing that the solution is acid. In the following section we are going to learn some more about the acid properties of carbon dioxide.

CARBONATES

Since carbon is a non-metallic element its oxide, carbon dioxide, forms an acid when combined with water. This acid is called carbonic acid, and it is made very simply by dissolving carbon dioxide in water.

$$CO_2 + H_2O \rightleftarrows H_2CO_3$$

We write the reaction with a double arrow because carbonic acid decomposes very easily, and in any solution of carbon dioxide, molecules of carbonic acid are continually forming and breaking down, so that there is a balance or equilibrium, between the forward and the reverse reactions.

Although carbonic acid is an extremely weak acid, that is, it produces only a few hydrogen ions, it is able to turn blue litmus paper red and to combine with many bases to form salts.

EXPERIMENT 255—Carbonic Acid.

Into a test tube half full of water drop a very small strip of red Litmus Paper (No. 18) and also a strip of blue Litmus Paper (No. 18).

Put about half a teaspoonful of baking soda or powdered washing soda into another test tube, add a small quantity of water and a few drops of vinegar. Immediately insert the gas delivery tube and stopper into this test tube and then put the end of the gas delivery tube in the test tube containing the 2 pieces of litmus paper. Carbon dioxide gas will be given off from the soda and vinegar and will pass through the delivery tube and bubble into the water containing the litmus paper. Keep a flow of gas for several minutes, occasionally adding 2 or 3 drops of vinegar to the test tube that contains the soda. Soon you will notice that the blue litmus paper is changing over to the same color as the red litmus paper, indicating that the carbon dioxide gas has dissolved in the water to form carbonic acid.

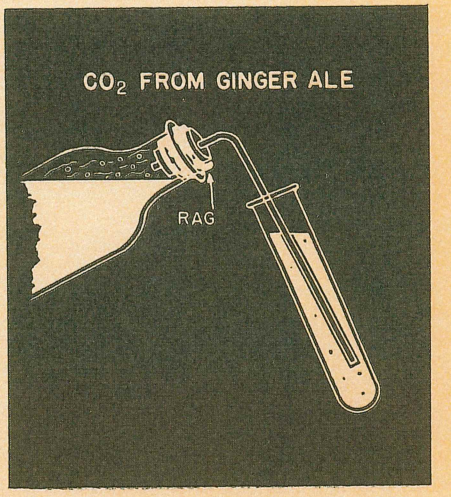
When the change from blue to red is completed, warm the test tube of carbonic acid gently and notice the large number of gas bubbles which form on the sides of the tube. Carbon dioxide is much less soluble in water at higher temperatures, so that some of it must escape as the solution is warmed. Boil the liquid for a few moments, taking care that it does not spurt out of the tube. Boiling decomposes the carbonic acid and drives carbon dioxide out of the solution, so that now if you drop in a piece of blue litmus paper, it will remain blue, showing that an acid is no longer present.

The color changes in this experiment will depend a good deal upon the impurities, such as hardness, in the water which you use. Rain water or distilled water will work best.

Carbonated beverages, such as gingerale and soda water, are saturated with carbon dioxide under pressure. When the pressure is released, so much gas cannot remain dissolved, and so it comes out in the form of bubbles. Soda water is so-called because it was formerly made with carbon dioxide that had been prepared from soda.

EXPERIMENT 256—Identifying the Gas from Gingerale.

Wrap a narrow strip of cloth or paper around the cork of your gas delivery tube so that it will fit the mouth of a bottle of gingerale or other carbonated beverage. Fit your delivery tube into the mouth of a newly opened bottle and place the other end of the tube in a test tube ¾ full of clear lime water, made by adding half a measure of Calcium Oxide (No. 20) to a test tube ¾ full of water, shaking thoroughly, allowing to settle, and then pouring off the clear lime water into another test tube. Allow the gas to bubble through the lime water for a few minutes. The white precipitate of calcium carbonate proves that the gas is carbon dioxide.



The carbonates of many metals are insoluble in water, so they can be precipitated from solutions by means of a soluble carbonate, such as sodium carbonate. Since carbonic acid is a very weak acid, solutions of soluble carbonates are usually alkaline and turn red litmus paper blue.

EXPERIMENT 257—A Test for Carbonates.

Suppose you have some substance which you believe is a carbonate. How can you test it? Since all carbonates give off carbon dioxide gas when treated with acid, you need only to add acid to your unknown substance and test it with lime water to see whether the gas, if any, which comes off is carbon dioxide.

Prepare some clear lime water by stirring 3 measures of Calcium Oxide (No. 20) in a glass half full of water for a minute or two. Let the undissolved solid settle to the bottom and fill a clean test tube ¼ full of the clear liquid.

In another test tube place 4 measures of the substance you want to test. It should be crushed or ground to small pieces. Add to it 5 measures of Sodium Bisulfate (No. 7) and a few drops of water, and then fit the test tube quickly with the gas delivery tube and stopper. Dip the other end of the gas delivery tube into the lime water in the other test tube. If carbon dioxide gas is being set free, it will make the lime water cloudy, and prove that your unknown substance contains a carbonate.

Test in this manner any of the following substances:

Baking soda (sodium bicarbonate), washing soda (sodium carbonate), baking powder (a mixture of sodium bicarbonate with potassium acid tartrate, calcium acid phosphate or alum), oyster shells (which consist chiefly of calcium carbonate), chalk, tooth powder, and plaster (which also contains calcium carbonate).

EXPERIMENT 258—Making Sodium Carbonate from Sodium Bicarbonate.

Dissolve 1 measure of baking soda (sodium bicarbonate) in a test tube half full of water and add 1 or 2 drops of Phenolphthalein Solution (No. 22). Unless your water supply is quite alkaline, the solution will be colorless.

Now warm the test tube over a flame and notice how much carbon dioxide comes off in the form of bubbles even before the liquid boils. Boil carefully for a few moments, and the solution will slowly turn pink, showing that it has become alkaline due to the formation of sodium carbonate.

$$\underset{\text{(sodium bicarbonate)}}{2\ NaHCO_3} + \text{Heat} \rightarrow \underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O}$$

EXPERIMENT 259—Making Sodium Bicarbonate from Sodium Carbonate.

The reverse of the preceding experiment is to add enough carbon dioxide to sodium carbonate to convert it into the bicarbonate. The equation for the change is just the reverse of the one we wrote above.

Dissolve ¼ measure of Sodium Carbonate (No. 4) in a test tube half full of water and add 1 drop of Phenolphthalein Solution (No. 22), coloring the solution red.

Now generate CO_2 gas with sodium bicarbonate and vinegar and pass carbon dioxide through the sodium carbonate solution for several minutes. Slowly the pink color of the phenolphthalein begins to fade as the sodium carbonate absorbs carbon dioxide and changes into sodium bicarbonate. The final loss of color comes rather suddenly.

$$Na_2CO_3 + CO_2 + H_2O \rightarrow 2\ NaHCO_3$$

EXPERIMENT 260—Decomposition of Dry Sodium Bicarbonate.

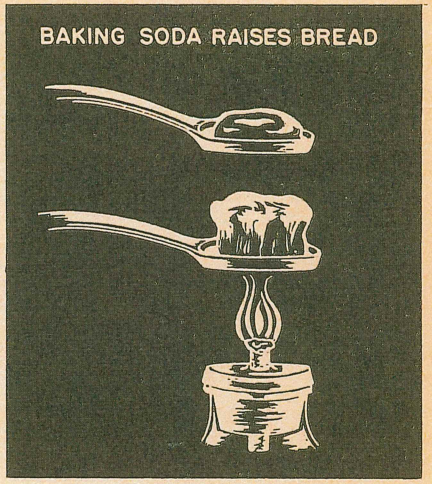
Fill a test tube half full of water and add 1 drop of Phenolphthalein Solution (No. 22). Unless your water supply is too alkaline, there will be no color. Drop in ½ measure of baking soda and notice that the solution remains colorless.

Now place several measures of baking soda in your spoon and heat it strongly for several minutes over a flame. Allow the substance to cool and then drop it into the water containing phenolphthalein. Shake the test tube and the mixture will turn pink, showing that some of the dry sodium bicarbonate has been converted to sodium carbonate.

EXPERIMENT 261—How Baking Soda Raises Bread.

Ordinary baking soda is simply sodium bicarbonate, and as you saw in the preceding experiment it gives off carbon dioxide gas when heated. Therefore, when it is mixed into biscuit or cake dough, it forms gas bubbles throughout the dough during baking, producing a porous light texture. Weakly acid substances, such as sour milk or cream of tartar, are often added to assist in the liberation of carbon dioxide.

Put a teaspoonful of flour into a cup and add 4 measures of baking soda. Stir water into the flour in small portions until you have a good smooth dough. Place a piece of this dough about the size of a marble in your spoon and heat it gently over the alcohol lamp flame. Notice how it puffs up and becomes porous due to the gas which forms inside.



When lime is used in mortar, it slowly absorbs carbon dioxide from the air and is converted into calcium carbonate, just as carbon dioxide forms a precipitate in lime water. Calcium carbonate is the chief constituent of limestone and marble, and when it is formed from the lime in the mortar, it gives strength to the masonry. If we expose some lime water to the air, calcium carbonate will soon be formed.

EXPERIMENT 262—Limestone Ice.

Make some clear lime water by stirring 2 measures of Calcium Oxide (No. 20) in half a glass of water for a minute or two. Let the suspended solid settle to the bottom and pour off the clear solution into a saucer. Let it stand in a quiet place overnight, and in the morning you will find that a thin white crust has formed over the lime water. Although only 0.04% of the air is carbon dioxide, the calcium hydroxide solution has absorbed enough of it to make this film of carbonate. Add a drop or two of vinegar and watch the crust dissolve, giving its carbon dioxide back to the air.

EXPERIMENT 263—Strontium Carbonate.

Dissolve 1 measure of Strontium Chloride (No. 10) in a test tube half full of water and add a few drops of sodium carbonate solution prepared by dissolving 3 measures of Sodium Carbonate (No. 4) in a test tube half full of water. The white precipitate is strontium carbonate. Strontium is an element which closely resembles calcium in its properties, so it is not surprising to find that their carbonates are very similar. Strontium carbonate is found in nature as the mineral strontianite.

The carbonates of some metals are hydrolyzed to so great an extent that they cannot exist in contact with water. When you try to make them by the following method you obtain a precipitate which is a sort of cross between the carbonate of the metal and the base, or hydroxide, of the metal, that is, you obtain a basic carbonate. The exact composition of such a compound often depends upon the temperature and concentration of the solutions from which it is precipitated.

EXPERIMENT 264—Basic Magnesium Carbonate.

Add 1 measure of Epsom salts (magnesium sulfate) to a test tube half full of water and dissolve the solid by shaking the tube. Now add a few drops of sodium carbonate solution prepared by dissolving 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water, and a white precipitate of basic magnesium carbonate will be formed. This is somewhat similar to the mineral magnesite.

EXPERIMENT 265—Basic Cobalt Carbonate.

Repeat the preceding experiment using about 8 drops of Cobalt Chloride Solution (No. 27) in place of the magnesium sulfate. When you add sodium carbonate solution, a light blue precipitate of basic cobalt carbonate is formed.

EXPERIMENT 266—Basic Copper Carbonate.

Place 1 measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7) in a test tube. Fill the tube half full of water and shake it for a few moments until the solution becomes clear. Now add a few drops of sodium carbonate solution prepared as before, and a very pretty blue precipitate of basic copper carbonate will be formed. The minerals malachite and azurite are basic carbonates of copper.

EXPERIMENT 267—Basic Zinc Carbonate.

Put a small piece of zinc about the size of the end of the measure into a test tube and add 2 measures of Sodium Bisulfate (No. 7). Fill the test tube ¼ full of water and warm it gently over a flame for a minute or two. Pour off the clear solution into another test tube, and add to it a few drops of sodium carbonate solution prepared as in the preceding experiments. The white precipitate is basic zinc carbonate.

A few metals refuse to form even basic carbonates, and when sodium carbonate is added to solutions of their salts, the base, or hydroxide, of the metal is formed instead of the carbonate or basic carbonate.

EXPERIMENT 268—Precipitation of Ferric Hydroxide by Sodium Carbonate.

Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water, and add a little sodium carbonate as in the preceding experiment. A dense brown precipitate forms at once and carbon dioxide gas is given off. It is just as if ferric carbonate first formed but then immediately hydrolyzed, giving ferric hydroxide and carbonic acid.

EXPERIMENT 269—Precipitation of Aluminum Hydroxide by Sodium Carbonate.

Dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube half full of water, add a little sodium carbonate as in the preceding experiment. A dense white precipitate of aluminum forms at once and carbon dioxide gas is given off. Aluminum carbonate, like ferric carbonate, immediately hydrolyzes, giving aluminum hydroxide and carbonic acid.

FUELS

If mankind had never learned how to use fuel, there would be no civilized human race. Long before the beginning of history, our primitive ancestors were using fire to keep them warm and to cook their food. Today we are still using fire for those two purposes, but we have learned to use it also in many other ways. Fuel burned under boilers gives us steam to turn the wheels of industry, to drive ships across the sea, and to generate electricity, the bearer of light and power. Gasoline and fuel oil burned in the cylinders of engines propels us along the highways, the railways and through the air. In blast furnaces and smelters, it is the burning of fuel which unlocks metals from the ores in which they are hidden.

The most important fuels of our modern times are coal, petroleum and natural gas. All of these fuels have been formed from the remains of prehistoric vegetation, which became covered with millions of tons of earth and rock. During the ages, the heat and pressure gradually converted this buried vegetation into carbon (coal) and hydrocarbons (petroleum and natural gas).

Let us begin the study of fuels by seeing what we can learn from a candle. Candles are usually made from paraffin, a substance obtained from petroleum,

and generally have some stearic acid added to harden them. Paraffin is a mixture of compounds composed of the elements carbon and hydrogen. Such compounds made up only of carbon and hydrogen are called hydrocarbons. When a hydrocarbon burns completely, the carbon is converted into carbon dioxide, and the hydrogen into water. We learned, when we studied water, that when a candle burns, the hydrogen contained in the candle combines with the oxygen of the air, forming water. Now let us prove that carbon dioxide is also formed when a candle burns.

EXPERIMENT 270—Showing that Carbon Dioxide Is Formed When a Candle Burns.

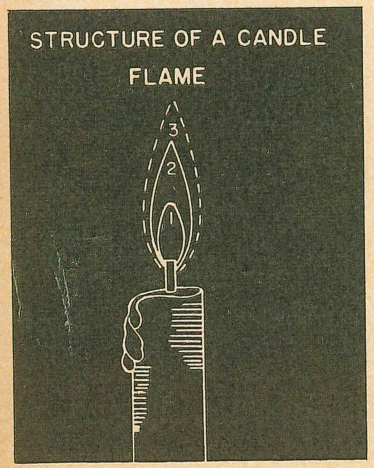
Place 2 measures of Calcium Oxide (No. 20) in half a glass of water and stir the mixture for a minute or two. Let the solution stand until the powder has settled to the bottom and then pour off the clear liquid into another glass. This liquid is lime water.

Set a candle in the middle of an ordinary sauce or baking pan, and stick it tight with a few drops of molten wax. Pour about 2 inches of water into the pan, light the candle and set an inverted bottle or fruit jar over it.

When the candle flame has gone out, remove the jar or bottle from the water, turning it right side up as soon as possible to keep the heavy gas from spilling out, and quickly pour in a little of the lime water. Put your hand over the mouth of the bottle or jar and shake it for a moment. If you will look closely you will see that the lime water is now turbid with white particles. As we have already learned, the formation of this white precipitate of calcium carbonate is a test for carbon dioxide.

A flame is simply a body of burning gas. In order to burn with a flame, a substance must first be vaporized, or converted into a gas. A candle burns with a flame because the wax, melted by the heat, climbs up the wick, and is converted into vapor, which then combines with the oxygen of the air so rapidly that it gives off heat and light.

EXPERIMENT 271—The Structure of a Candle Flame.



Look closely at a candle flame and notice its two main parts. First, the dark zone around the wick, and second, the bright yellow luminous part of the flame. Outside this is a third zone, consisting of the invisible, hot products of combustion.

The dark inner zone consists of unburned gas given off from the wick of the candle. Since there is no air at the center of the flame, this gas cannot burn at first, and so this portion of the flame is relatively cool, as you can show by thrusting a match stick into the center of the flame just above the wick. The part of the stick held in the dark zone does not char as soon as the part passing through the side of the flame.

Candle wicks are woven in such a way that as the candle burns, the end of the wick bends over into the hot outer part of the flame. This keeps it burned off short and prevents the flame from smoking.

EXPERIMENT 272—Removing Gas from a Candle Flame.

With care, a portion of the unburned gas can be drawn off from the center of the flame. Hold one end of a glass tube in the flame directly over the wick. Let the tube slant upwards so that the hot gas will rise through it and issue from the other end, where it can be lighted with a match.

EXPERIMENT 273—Igniting Candle Gas.

Blow out a candle with a long wick that is burning and after an instant relight the smoke at a little distance from the end of the wick. Notice how much easier it is to light a candle just after it has been extinguished than after it becomes cold. Gas is still coming off from the hot wick and is ready to be ignited, but in lighting a cold candle, wax must be melted and vaporized before it can burn. In an alcohol or kerosene lamp, as well as in a candle, the fuel must be vaporized before it can burn.

EXPERIMENT 274—The Luminous Zone—Making Lampblack.

The second, or bright yellow part, of the flame contains particles of white-hot carbon, formed by the action of heat on the gas of the inner zone. These particles burn completely in the outside of the flame. Their presence can be proved by holding a cold spoon or piece of glass in the flame for a few moments. It will become covered with a black deposit of soot, or lampblack. The cold object chills the flame and prevents the carbon from being completely burned.

In making lampblack on a large scale, natural gas is burned in large chambers and the flame is chilled by iron pipes through which water circulates. Large quantities of lampblack are used in printing ink and as a toughening agent in rubber.

WOOD

Wood is still used as a fuel in some localities. This important material contains carbon, hydrogen, oxygen and nitrogen in the form of complex compounds. When wood is burned with a plentiful supply of air, we would naturally expect to form carbon dioxide and water with smaller amounts of other substances, which are found largely in the ash or non-combustible part of the wood.

EXPERIMENT 275—Testing for Carbon Dioxide in the Combustion Products of Wood.

Light a splinter of wood, somewhat larger than a match stick, and when it is burning brightly drop it into a milk bottle or fruit jar, and cover the top of the jar with a piece of cardboard.

When the flame has gone out, pour some clear lime water into the jar and shake it around. In this case, you will find that the lime water turns white or turbid, showing that carbon dioxide is formed when wood burns.

EXPERIMENT 276—Testing for Carbon Dioxide in the Combustion Products of Paper.

Since most paper is composed largely of wood fibers, we would expect to find the same products of combustion from burning paper as we do from burning wood. Light a small strip of paper and drop it into a jar or milk bottle, covering the top of the jar with a piece of cardboard. When the flame has gone out, test for carbon dioxide gas by pouring in a little clear lime water and shaking the jar.

EXPERIMENT 277—Showing that Water Is Formed When Wood Burns.

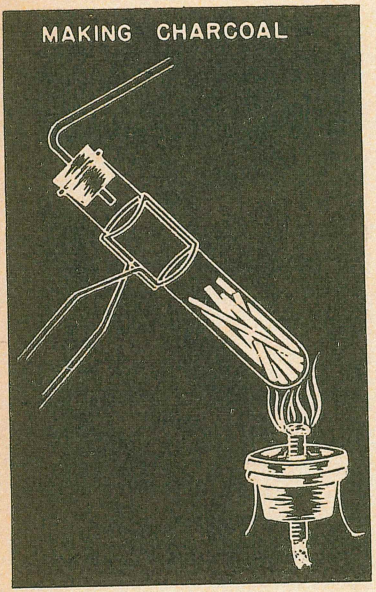
Light a small splint of wood and as soon as it is burning well, hold it upright and invert a dry, cool glass tumbler over the flame. After a few seconds remove the tumbler and notice that moisture is formed on the inside surface of the glass. This proves that water is formed when wood burns.

EXPERIMENT 278—Showing that Water is Formed When Paper Burns.

Roll a piece of paper, about 4 inches square, into a loose tube. Light one end of this tube and as soon as the paper is burning well, hold it in an upright position and invert a dry, cool glass tumbler over the flame. After a few seconds remove the tumbler and again examine it for moisture on the inside surface.

We have learned that when wood is burned with a plentiful supply of air (oxygen), carbon dioxide and water are formed. When wood is heated, however, in a closed vessel or retort so that the oxygen of the air cannot unite with it, a complicated chemical decomposition occurs and a

variety of compounds is obtained. A combustible mixture of gases are given off from the wood, also vapors of various liquids, wood alcohol, acetic acid (wood vinegar), some tar-like substances and moisture. The residue is charcoal, composed almost entirely of carbon, along with the mineral matter (ash) of the original wood.



EXPERIMENT 279—Making Charcoal.

Break the heads off 5 or 6 wooden matches and put the sticks in the bottom of a test tube. Now heat the tube, using your test tube holder, over an alcohol lamp or gas flame.

When smoke begins to pass from the test tube, hold a piece of moistened blue Litmus Paper (No. 18) in the smoke. Notice that it is turned red, showing that there is acid in the gas. Now fit the tube with your gas delivery tube and continue heating. You will find that the gas which issues from the tube is inflammable, and after the distillation has continued for a few minutes you can light it as it comes out of the tube.

Continue heating until gas no longer comes off, then let the test tube cool and empty out the charcoal which remains, using your stirring rod to loosen it from the sides, if necessary. Charcoal is an impure variety of amorphous carbon, that is, carbon which is not crystalline.

COAL

When coal is heated in a closed vessel, coal gas is evolved and also ammonia, along with a great variety of oily and tar-like liquids. The distillation or carbonization of coal is an important source of ammonia for use in fertilizers, and from the oil and the coal tar may be prepared the raw materials of dyes, drugs, disinfectants, high explosives, perfumes, and a multitude of other chemical products. The residue left behind in the retorts is coke, composed, like charcoal, of amorphous carbon. Its chief use is in blast furnaces and smelters for obtaining metals from their ores.

Although a large fraction of our coal production is heated in coke ovens to obtain coke along with the valuable by-products we have mentioned, most of it is still burned directly, wasting millions of tons of ammonia and precious organic chemicals. Coke is clean to handle and smokeless to burn, and in the future coal will probably be more and more converted into coke before being burned, the by-products—gas, ammonia, oils, and tar—being collected and used in industry.

EXPERIMENT 280—Making Coke.

Put 4 measures of powdered soft coal into a test tube and heat it over a flame. Soon you will notice fumes coming from the mouth of the test tube. Hold a piece of cold glass or porcelain at the mouth of the tube as you heat it, and a yellow or brown liquid will condense upon it. This is similar to crude coal tar. Fit your gas delivery tube to the test tube, and when gas begins to issue from it, light it and it will burn. When purified, this gas is the same as the coal gas which is often used for heating.

When smoke no longer issues from the coal, stop heating, and after the tube cools, examine the substance which remains. You will find that it is coke.

OIL FROM THE EARTH

Petroleum is an oily liquid found below the surface of the earth at many places. Usually it is a dark, thick liquid, often with an unpleasant odor, but different samples of petroleum vary greatly in their appearance and properties. Sometimes it is almost as clear and fluid as gasoline, while other specimens may be black and semi-solid, like asphalt. It is usually found in strata of porous rock or sand, sometimes near the surface, in other places miles down in the earth.

Petroleum consists principally of a large assortment of hydrocarbon compounds, with a small percentage of nitrogen and sulfur compounds. Samples coming from different places may have very different composition, and not all are equally suited for all purposes. Most petroleum is supposed to have been formed by the decomposition of plant and animal remains which long ago were buried deep in the earth.

In refining petroleum, it is boiled in a still, and the vapors are led through a high tower in which they are condensed back to liquid. The heavy oil with a high boiling point condenses first near the bottom of the tower; then higher up, the lighter oils and kerosene; then naphtha, used as a solvent for rubber and for lacquers; and finally, the light, volatile gasoline, having the lowest boiling point of all. Each of these products must be further purified before it is ready for market. The gasoline and kerosene are washed with alkali and sulfuric acid to remove sulfur compounds and other undesirable substances. The oils are chilled to —40° F., at which temperature paraffin wax freezes out as a solid. From some of the heavier oils, vaseline, or petrolatum, is separated by a similar procedure. Petroleum from some wells, especially in California and Texas, leaves a residue of asphalt in the still after the volatile constituents have boiled off.

Since 1920 the rapid growth in the number of automobiles has created an ever increasing demand for gasoline, a demand which could not possibly be satisfied by the natural gasoline in petroleum. A large amount of our gasoline at present is made from the heavier, less volatile portions of the petroleum by a process called *cracking*. Kerosene and various grades of oil are heated to about 950° F. in a coil of pipe under pressure of approximately 400 pounds per square inch. The vaporized oil passes into a huge steel chamber, where the large hydrocarbon molecules crack apart into smaller fragments, more suitable for motor fuel. Cracked gasoline contains a different type of hydrocarbon compound from that which makes up natural gasoline, and this makes it burn more smoothly without knocking in modern high-compression engines. Cracking, therefore, produces not only more gasoline from a given amount of petroleum, but also better gasoline.

Further improvement in the quality of gasoline results from adding small amounts of certain compounds to it. The most commonly used is lead tetraethyl, $Pb(C_2H_5)_4$, which has a very marked effect in reducing knocking in high-compression engines.

Diesel engines are now used to pull many trains and for many passenger ships, and smaller models are being developed for automobiles and airplanes. The big advantage of Diesel engines is that instead of gasoline, they use heavy, comparatively non-volatile oil. They are, therefore, much cheaper to operate than ordinary gasoline engines. Since their fuel cannot be vaporized in a carburetor and sucked into the cylinder, it is sprayed in under pressure. The hot, compressed gases in the cylinder cause it to ignite, so that no spark plug is needed.

One of the things which oil chemists and engineers want to know about oils is the temperature at which they will take fire. If this temperature, called the flash point, is too low, there is danger of fire or explosion when the oil is used to lubricate machinery. The flash point of kerosene is important when it is used in a lamp or a stove.

To determine the flash point of an oil, it is heated slowly in a metal cup, and at intervals, a small flame is passed across it at a certain height above its surface. Its temperature is read on a thermometer when it first takes fire with a flash as the flame is passed over it.

EXPERIMENT 281—Flash Point of an Oil.

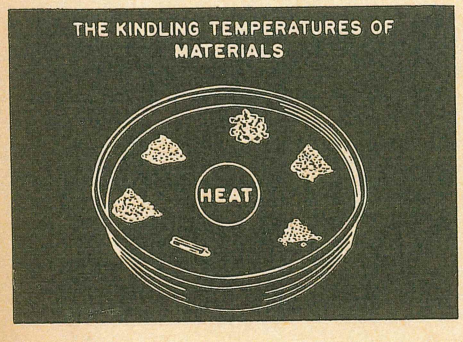
Pour a few drops of lubricating oil into your spoon and warm it slowly over a flame. Soon fumes will begin to be evolved. Test them from time to time by bringing the edge of the spoon into the flame. When the flash point of the oil is reached, the fumes will take fire. Try this experiment with automobile crankcase oil, with light machine oil, and with a few drops of kerosene, and notice the differences among them.

EXPERIMENT 282—Flash Point of Alcohol.

Put a few drops of alcohol which you use for your lamp into the spoon and heat it gently over a flame. Notice how much sooner it takes fire than the oil in the preceding experiment. Because of their low flash point, rubbing alcohol and radiator alcohol should be used with care when open flames are near.

Just as oils and other inflammable liquids catch fire at different temperatures, so solid substances differ greatly in the ease with which they can be set on fire. The temperature to which a substance must be heated before it begins to burn is called its kindling point, or ignition temperature.

EXPERIMENT 283—Ignition Temperature.



Obtain a tin plate and arrange in a circle, about 2 inches in diameter, all of the following substances which you can obtain. Put each substance in a little pile, so that there will be about an inch between piles. Use 2 measures of powdered coal, 2 measures of powdered charcoal, a small piece of paper crumpled up into a ball, a short piece of match stick, the head of a safety match, 2 measures of Sulfur (No. 1) and 2 measures of sugar.

Put the pan on a support and set your alcohol lamp under the center so that the flame will strike the pan exactly in the center of the circle, an equal distance from each material.

Watch closely and see which substance catches fire first, and which is the last to burn.

EXPERIMENT 284—Lowering the Ignition Point of Sugar.

Holding a lump of sugar in your test tube holder, bring one end of it into a candle flame and try to light it. Although the sugar turns black and melts, you will find it very difficult to make it burn with a flame.

Now touch the end of the lump to some cigarette ashes so that a small amount of the ash sticks to the sugar. Again hold the sugar in the flame, and this time you will find that it burns easily and that it will continue to burn even after you remove it from the flame. The ash contains small quantities of metallic oxides which act as a catalyst and promote the combustion of the sugar. Since the tobacco ash does not itself burn but nevertheless promotes the combustion of the sugar, it is an example of a catalyst. A catalyst is a substance which increases the speed of a chemical reaction without itself being permanently changed. Instead of cigarette ash, numerous other substances, such as iron rust or sodium carbonate, may be used in this experiment.

MATCHES

The world uses a greater number of matches than of any other manufactured article, yet the cheap, convenient, and safe modern match is a rather recent invention. The history of man's methods of making fire, from the cave man rubbing two sticks together down to the present scratching of a match, is extremely interesting, and many ingenious methods have been invented, used for a while and then discarded in favor of something better.

Modern matches are made by automatic machines which turn them out at the rate of millions a day. In the case of wood matches, an automatic cutter slices the blocks and inserts the match sticks into holes in an endless belt on which they pass through the dipping machine. In making paper matches, the cardboard is diecut and long strips are fed through the dipping machine, after which they are cut to correct sizes.

The strike-anywhere match contains phosphorus sesquisulfide, P_4S_3. The head of a safety match contains no phosphorus, but only a mixture of an oxidizing agent with an easily combustible material. The striking panel on the side of the box or on the folder which holds the matches, however, is coated with a mixture which contains red phosphorus and friction upon this panel ignites the head of the match.

EXPERIMENT 285—Proving that Matches Contain Sulfur.

Cut off the heads of several matches. Place these in a test tube and add 3 or 4 drops of water, and upon shaking the tube, a thin paste will be formed. Place a drop or two of this paste on a shiny piece of silverware or a silver coin. Heat the silver over your alcohol lamp until the pasty material ignites. After this paste burns, notice the black stain which appears on the silver. This is due to the formation of silver sulfide, formed by the sulfur in the matches uniting with the silver. The formation of black silver sulfide proves that the matches contain sulfur.

ORGANIC CHEMISTRY

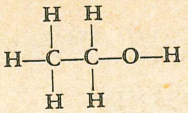
Many years ago chemists noticed that substances obtained from animal or vegetable life always contain carbon. Substances obtained from minerals on the other hand usually did not contain carbon, with the outstanding exception of the carbonates of the various metals. Based on these observations, the term organic chemistry was applied to compounds derived from substances of animal or vegetable origin. The term inorganic chemistry was applied to compounds derived from substances of mineral origin.

The fact that all substances derived from animal and vegetable life contain carbon led also to the belief that the carbon atom and life go hand in hand. Organic chemistry was, therefore, defined as pertaining to substances which contain carbon and are obtained from life. Later it was found that organic substances containing carbon could also be made in the laboratory without the aid of life. Hence, today we define organic chemistry as the chemistry of carbon compounds.

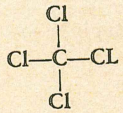
In inorganic chemistry, a majority of the reactions are based on ionization and these reactions occur as a rule almost instantaneously when the substances are brought together in the form of water solutions. In organic chemistry we find that most organic compounds are not soluble in water and are not readily ionized. As a rule organic reactions are slower and more frequently require the aid of heat as well as time.

The organic chemists are principally interested in the arrangement of atoms within the molecule, and this accounts for the practice of showing molecular structure in detail when writing organic formulas and equations. For example, the formula of ethyl alcohol, or grain alcohol, might be

written simply as C_2H_6O. It is more informative, however, to show how the atoms are arranged in the molecule by writing a structural formula:



Such a detailed structural formula is often abbreviated to CH_3CH_2OH or even to C_2H_5OH, at the sacrifice, however, of some clarity, at least for the beginner. Methanol, or methyl alcohol, has the formula CH_3OH. See if you can figure out how the complete structural formula of this compound would look. Carbon tetrachloride is CCl_4. The structural formula, therefore, would be written:



The writing of organic formulas on paper can be likened to the drawing of an architect's plans for a building. Today the organic chemist can even predict in advance the probable properties of new organic compounds before they are made in the laboratory.

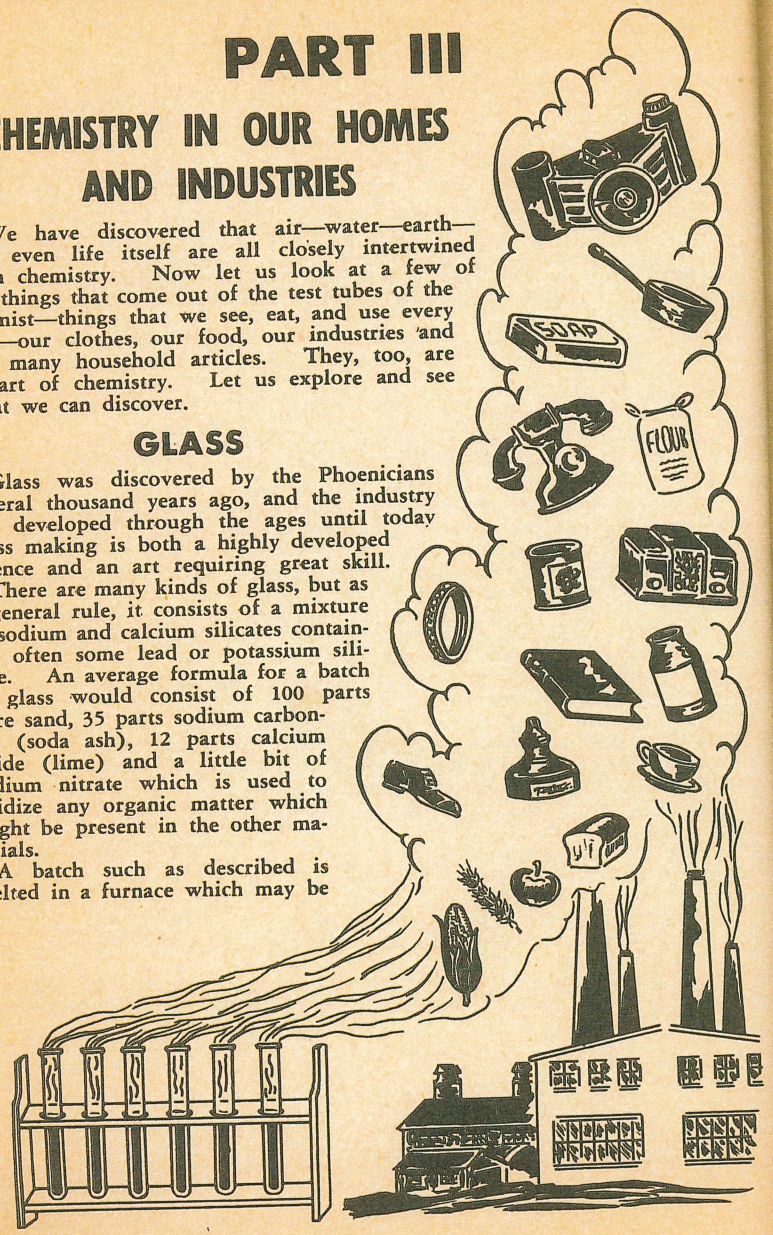
Almost all modern dyes, drugs, and explosives are organic compounds. A few familiar examples are indigo, sulfanilamide, procaine, thiamine, ascorbic acid, penicillin, TNT, and nitroglycerine. Foods, including fats, carbohydrates, proteins, vitamins, and flavors, are organic; rubber, natural and synthetic, resins, plastics, fuels, paints, paper, wood, and clothing are composed of organic substances. The life of all plants and animals is absolutely dependent upon reactions of organic compounds. Thus we see that all of us come in contact with organic chemistry a thousand times a day.

In the following part of this book, "Chemistry in Our Homes and Industries," you will take up several practical applications of chemistry which are ordinarily classified as organic chemistry.

PART III

CHEMISTRY IN OUR HOMES AND INDUSTRIES

We have discovered that air—water—earth—and even life itself are all closely intertwined with chemistry. Now let us look at a few of the things that come out of the test tubes of the chemist—things that we see, eat, and use every day—our clothes, our food, our industries and our many household articles. They, too, are a part of chemistry. Let us explore and see what we can discover.



GLASS

Glass was discovered by the Phoenicians several thousand years ago, and the industry has developed through the ages until today glass making is both a highly developed science and an art requiring great skill.

There are many kinds of glass, but as a general rule, it consists of a mixture of sodium and calcium silicates containing often some lead or potassium silicate. An average formula for a batch of glass would consist of 100 parts pure sand, 35 parts sodium carbonate (soda ash), 12 parts calcium oxide (lime) and a little bit of sodium nitrate which is used to oxidize any organic matter which might be present in the other materials.

A batch such as described is melted in a furnace which may be

either a clay pot that is set in a hot furnace, or a tank furnace, which is merely a tank built directly in a furnace made of heat-resistant bricks. The batch of material is left in the hot furnace for 24 hours or more until it melts and becomes fluid and free from bubbles. This melted glass is then ready to work and form into the desired shapes.

EXPERIMENT 286—Making Soft Glass.

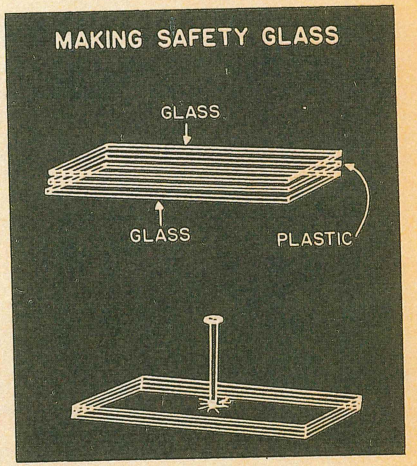
Put 5 drops of Sodium Silicate Solution (No. 19) in your spoon and mix thoroughly with 1 measure of Sodium Carbonate (No. 4) and ⅓ measure of Calcium Oxide (No. 20). Hold your spoon in the flame of your alcohol lamp. You will notice that when you heat the mixture it will swell, turn white, froth, and finally assume a shiny glass-like appearance. This is soft glass, which is a soda-lime silicate. If you find that the heat of your alcohol lamp flame is not sufficient to melt the glass, hold your spoon over the flame of a gas burner.

To make plate glass, the molten glass is poured on a large, flat table and rolled to the desired thickness by a metal roller. The plate glass is then put in an oven and annealed, or allowed to cool very slowly. It is then smoothed and polished.

Wire glass is plate glass into which wire has been imbedded during the rolling process. Safety glass which will not shatter is made by cementing a sheet of transparent plastic between two pieces of plate glass.

EXPERIMENT 287—Making Safety Glass.

You can make a piece of safety glass from two microscope slides. Cover one side of each with Canada balsam or cellulose acetate cement, such as is used in gluing model airplanes. Then cut a piece of clear plastic and place this between the coated slides. Press the two pieces of glass and the plastic together and place them in a clamp to dry. In two or three days, after the cement has dried, you will notice that you can see through the two pieces of glass and plastic; just as well as if you were looking through only one piece of glass.



EXPERIMENT 288—Testing Safety Glass.

Test the piece of "safety" glass which you made in the preceding experiment by driving a nail through it and note that it does not shatter, but holds together and merely cracks. Keep your face at a distance as you test the glass.

Window glass is now made by machinery but formerly it was made entirely by hand. The workman would gather a lump of molten glass on the end of a hollow iron pipe and by alternately blowing and swinging he produced a large hollow cylinder of glass. The cylinder was allowed to harden and then was cut lengthways along one side. The next step consisted of putting this cut cylinder into an oven and gradually heating it until the glass began to soften and the cylinder flattened out into a pane. Machine-made window glass is made very much the same way, except that the operation of dipping the hollow pipe into the molten glass and blowing are done mechanically.

Bottles and similar heavy glassware are made by dropping some molten glass into a mold and forcing the glass to take the shape of the mold by

means of a plunger or compressed air, or both. On modern bottle making machines each operation is automatic and bottles are turned out at an amazing rate of speed.

Almost without exception, articles made of glass have to be cooled very slowly, or annealed, as the process is called. If glass is cooled suddenly, the outside will harden while the interior of the glass is still in a semi-liquid state and under these conditions great strains develop so that the glass is likely to fall to pieces due to a change in temperature or a scratch on the surface. To obtain glass free from strain, the finished article, while it is still hot, is placed in a heated room or oven and allowed to cool very slowly. The time of cooling or annealing varies from several hours to a week or more, depending on the thickness and composition of the glass.

An interesting recent development in the glass industry is a strong comparatively flexible glass which is very resistant to sudden temperature changes and to shock. A piece of it ¼ inch thick will not crack when molten lead is poured over it while it rests on a block of ice. When it is broken by a blow, the fragments are rounded instead of sharp and jagged. It is made by a carefully controlled chilling of the glass from a high temperature in such a way that the interior contracts, compressing the surface layers.

Many special kinds of glass are made; for example, cut glass, which has a brilliant, shiny appearance. It is made by adding lead oxide to the batch before melting. The designs or decorations are put on by means of grinding wheels after the glass has cooled.

Glass made with a high percentage of silica, some boric oxide, very little soda, and no lime is more resistant to heating and cooling and to chemical reagents. Such borosilicate glass is used under various trade names for kitchenware and for laboratory equipment. Articles made of it do not expand on heating nor contract on cooling nearly so much as if made of ordinary glass. For this reason, great strains are not set up in them as they are heated and cooled, and they are much less likely to break.

EXPERIMENT 289—Why Glass Cracks When Suddenly Chilled.

Heat a small piece of glass tubing in your alcohol lamp flame until it softens, and then plunge it quickly into water. The tube will crack due to unequal contraction on sudden cooling. Of course, when you plunge the hot tube into cold water the surface of the glass cools instantly, while the inside of the glass is still hot, and the unequal contraction which takes place when the glass is cooling causes it to crack.

Glass for special purposes may contain in addition to the usual constituents a great variety of others, for example, potassium, zinc, barium, lead, boron, phosphorus, arsenic, aluminum and antimony. Colored glass is made by adding to the molten batch suitable metal oxides or salts. The same substance may produce different colors, depending upon the amounts used and the method of adding it.

EXPERIMENT 290—Coloring Glass Light Green.

Make a small batch of soft glass in your spoon by heating intensely a mixture of sodium silicate, sodium carbonate and calcium oxide, as you did in the first experiment of this section, but add in addition a tiny crystal of Ferric Ammonium Sulfate (No. 21) about the size of a pinhead. You will find that your glass turns out to be greenish yellow.

If you look edgewise through a broken piece of cheap bottle glass, you will see that it is greenish. This is because of the iron contained as an impurity in the raw materials from which the glass was made.

EXPERIMENT 291—Coloring Glass Amber.

Make a glass mixture as before, adding a tiny amount of powdered charcoal. Notice the beautiful amber color which results.

EXPERIMENT 292—Coloring Glass Dark Blue.

Repeat the same process of glass-making, but this time use 1 drop of Cobalt Chloride (No. 27). Upon fusing this in your glass mix, you will have a glass of a dark blue color.

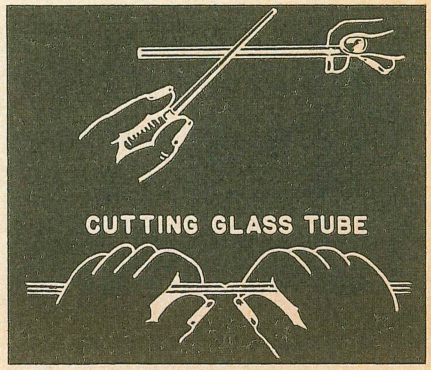
The value of glass as a building material is just beginning to be appreciated. Hollow glass blocks are used for the walls of windowless buildings, and beautiful decorative effects may be obtained, since the blocks can be made in any desired colors.

Fiber-glass is a fluffy, wool-like material which was developed for use in air filters for air conditioning installations. It is made by blowing jets of high-pressure steam against streams of melted glass. The individual glass fibers are only one-twentieth the diameter of a human hair, and a single pound of glass (the same amount used to make a beverage bottle) will make a single thread of more than 40,000 miles long. It is used in chemical plants for filtering corrosive liquids that would attack other filtering materials. Since it will not burn, it is valuable for insulating the walls of buildings against heat and cold. Its high strength and its resistance to fire suggest wide use in the manufacturing of rugs, curtains and electric wire insulation for reducing fire hazards in theatres, homes, factories, and ships at sea.

In working chemical experiments, it is often necessary to cut or bend glass tubing, and it is a great advantage to be able to do this well.

EXPERIMENT 293—How to Cut Glass Tubing.

Make a scratch with a three-cornered file in the place in which you wish to break the tube. Now hold the tube in your hands with your thumbs pressing upon the tube directly underneath the scratch, and then apply a pressure outward with your hands and break the tube just as you would break a stick. See the accompanying illustration. With a little practice, a clean and even fracture can be obtained.

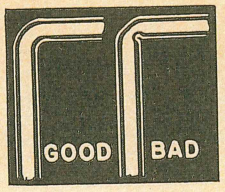


EXPERIMENT 294—Fire-Polishing Glass Tubing.

The cut ends of the piece of glass tubing have extremely sharp edges and are likely to cut stoppers and rubber tubes into which they are forced, to say nothing of your hands. It is good practice, therefore, to round off the sharp edges by fire-polishing. This is done by holding the end of the tube in the alcohol lamp flame until the glass softens and the edges become rounded. Observe that the hottest part of the flame is some distance above the wick. Be sure to rotate the tube slowly while you heat it so that it is heated evenly all around.

EXPERIMENT 295—How to Bend Glass Tubing.

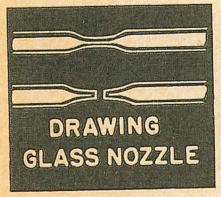
Place the part of the glass tube where the curve is desired in the flame of your alcohol lamp. The whole surface of the tube must be heated, so it is necessary to turn it in the flame to heat all sides. It is also best to use as wide a flame as possible so that the tube will be heated for a considerable length. This makes it easier to obtain a gradual bend and one that will not break easily. When the tube becomes hot and soft, remove it from the flame and bend it to the desired angle by simply applying a little pressure to the ends and bending it in much the same manner as you would a piece of wire.



Unless you heat the tube for a considerable distance your bend will kink and may not allow the proper passage of gases or liquids through the bent tube. This type of bend will break if slightly twisted. The illustration shows the RIGHT and WRONG way to bend glass tubing.

EXPERIMENT 296—How to Make Glass Nozzles.

A glass nozzle or tip is very convenient for use in burning small quantities of gas, as the gas coming out of a very small opening will burn better than when coming out of a large opening.



To make a glass nozzle, take a piece of tubing, at least 3 or 4 inches long and heat one part of it, turning it in the flame so as to heat all sides equally. When the glass begins to soften remove it from the flame and pull the tubing out, that is, pull on each end of the tubing and you will find that the place where the glass is heated will stretch out, becoming smaller and smaller until finally it is almost the size of a thread. When the tubing becomes cool, break off at any desired point, depending on the size of the opening which you want in the end of the nozzle. When breaking the thread, it is best to use a file scratch in order to be sure of a good, clean break. Of course, when you make a nozzle in this way you really have two nozzles, but you are quite likely to spoil one before you are through.

EXPERIMENT 297—How to Close the End of a Glass Tube.

Hold the end of the piece of tubing which you wish to seal or close in the hottest part of the alcohol lamp flame. Turn it between the thumb and fingers and continue to heat and you will see the end gradually coming together, until finally, the tube is entirely closed. Another way of closing a piece of glass tubing which is very much better if the tube is of large diameter, is to heat the tube all around as though you were going to bend it and when the glass becomes soft, pull from both ends and draw the glass out into a fine thread. Now break the two pieces apart and hold the end in the flame just long enough to form a nice round bottom where the glass has been pulled out.

EXPERIMENT 298—Joining Glass Tubes.

Take two pieces of tubing and heat the ends of each in the alcohol lamp flame at the same time until the edges start to round and curve inward. Then quickly press the two sections together. To make a smooth joint, it is necessary to reheat the place where the tubes are joined and work the glass together. This can be done best by closing one end of the joined tubes and blowing in the other very gently while the joint is hot and soft. By repeating this process a number of times you will find that the joint can be worked down until it is smooth, even, and strong.

EXPERIMENT 299—Etching Glass.

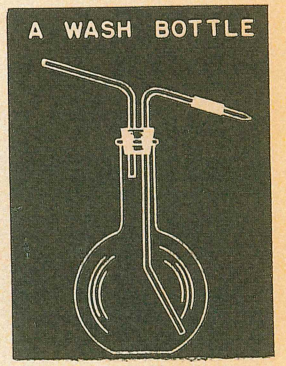
Obtain some very strong liquid glue and using a brush, paint a thick coat over the outside and inside of a tumbler. Set the glass away for about a day until the glue becomes dry and then put the tumbler in a warm place over a heater or near a stove, and allow it to remain there for about a day longer.

Now remove the glass from the warm place and put it where it is cool. About this time the glue will begin to crack and in doing so, will break off small chips of glass from the surface of the tumbler. After the glass has remained in a cool place for a little while, wash it with warm water to remove all traces of glue, and you will find very peculiar designs etched on the surface. If this experiment does not work properly with one coating of glue, apply two or more coats before putting it in a warm place. A very strong glue must be used.

EXPERIMENT 300—Frosting Glass.

Obtain a small quantity of Epsom salts, which is the same as magnesium sulfate. Dissolve 3 teaspoonfuls of Epsom salts in a glass of water and add about 1 teaspoonful of liquid glue, stirring the mixture thoroughly. Now apply this solution with a brush to the glass which you wish frosted. When the solution dries on the glass it will produce a splendid frosted appearance. This solution can be used to frost electric light bulbs, windows and many other articles.

EXPERIMENT 301—Making a Wash Bottle.



If you do not have running water in your laboratory, here is a piece of equipment which is easily made and will prove indispensable. To make a wash bottle, bend two pieces of glass tubing as shown in the illustration. You will also need a two-holed stopper, drawn glass nozzle as made in one of the experiments above, a small piece of rubber tubing and a Florence flask or bottle. Insert the two bent tubes through the stopper, first being sure that the ends are properly fire-polished, and attach the nozzle to the longer tube by means of the short piece of rubber tubing. Fill the flask ¾ full of water and place the cork assembly in the flask as shown. By blowing in the end of the shorter glass tube, you can direct a steady stream of water from the nozzle into test tubes, beakers or whatever you desire.

PAINTS, PIGMENTS AND COLORS

Since the dawn of history, the human race has been attracted by bright colors. Various kinds of paints were used thousands of years ago. Archeologists have uncovered Egyptian tombs in which were found beautiful colored paintings that have withstood the ravages of time. The American Indians used natural, mineral and vegetable colors to adorn themselves and their belongings. Today we are surrounded by colors on all sides and the use of color in our homes, factories and daily surroundings has almost become a science.

Coloring material used for protective and decorative purposes can be separated into two general classifications—paints, which consist of insoluble pigments mixed with a suitable vehicle or liquid—and stains, or dyes, which consist of a soluble color dissolved in a suitable liquid medium.

Paints can be further divided according to the vehicle, or liquid, used to carry the pigment. In general, these may be natural oils, such as linseed oil; synthetics, such as lacquers or liquid plastics; and finally water. Stains or dyes, may consist of organic solvents, such as turpentine, naphtha or alcohol, or the solvent may be water, depending upon the use of the stain.

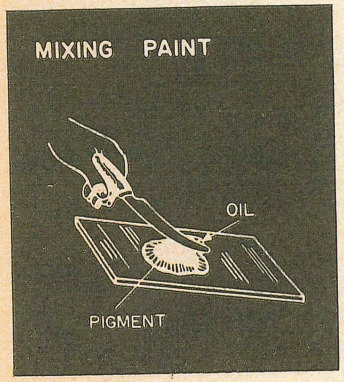
Oil paints are by far the most important of our coloring materials. These paints are used for outside work where protection of the surface is of importance, as well as the beauty of the color. Oil paints usually consist of linseed oil into which the pigment or color is very finely ground. Turpentine, or other thinner, is added to bring the paint to the proper consistency for brushing or spraying, and in some cases, dryers are added to hasten the drying process. Linseed oil, when exposed to the air, is oxidized, thus forming a hard, flexible, lasting film on almost any rigid surface,

such as wood, metal or brick. In the following experiments you will prepare some different colored pigments and grind them with linseed oil and turpentine to demonstrate the preparation of this class of paints.

EXPERIMENT 302—Making a Brown Pigment.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. In another test tube ⅓ full of water, dissolve 4 measures of Sodium Carbonate (No. 4). Pour the first solution into a second, shake the tube, and you will obtain a thick brown precipitate of ferric hydroxide. Let this precipitate settle, pour off the clear liquid and then scrape the precipitate into your spoon. Warm this precipitate in your spoon over a flame until it is thoroughly dried and no more moisture is given off. You now have brown oxide of iron, which is used as a pigment in brown paint.

EXPERIMENT 303—Brown Oil Paint.



Place the pigment which you prepared in the preceding experiment on a piece of clean glass, add a few drops of linseed oil and thoroughly mix the oil and the pigment with a flexible knife blade or with the back of a teaspoon. It will require quite a little mixing and rubbing to make a smooth oil paint. If the paint seems a little thick when you get it thoroughly mixed, add a few drops of turpentine, which will thin it so that it will brush out more smoothly.

EXPERIMENT 304—Making a Light Blue Oil Paint.

Place about 6 measures of Azurite (No. 39) on a glass plate, add a few drops of linseed oil and mix and grind this as before, adding a little turpentine, if necessary. This will make a very pretty light blue oil paint, sometimes called "mountain blue."

EXPERIMENT 305—Making Black Oil Paint.

See if you can obtain some lampblack or soot. If you or one of your neighbors uses fuel oil for heat, it should be easy to obtain a little soot from the inside of the furnace.

Place a small pile of soot on a glass slab, add a little linseed oil and again grind the mixture to the proper consistency. Lampblack is often used as a pigment in black oil paint.

Lacquers and synthetic enamels are often made from cellulose acetate or other synthetic plastic material dissolved in a suitable organic solvent or mixture of solvents. If no pigment is added, we have what is known as a clear lacquer. If a pigment is ground into the lacquer, we have the various colored lacquers with which we are familiar.

Synthetic plastic enamels are often made which require the application of heat to set them. This is not a simple drying, but brings about a change in the molecular structure of the plastic material. Automobiles are often finished with these materials.

EXPERIMENT 306—Making a Brown Lacquer.

Prepare some brown oxide pigment as in a previous experiment and when it is thoroughly dried, mix it on a glass slab with a few drops of clear liquid nail polish. If your lacquer is too thick when mixed, add a few drops of nail polish remover. Note that your lacquer, when applied to a surface, dries almost immediately. This is one of the big advantages of lacquer, as oil paints dry quite slowly due to the time required for the linseed oil to oxidize.

EXPERIMENT 307—Making Black Lacquer.

Proceed as in the above experiment, but use some lampblack instead of the brown iron oxide. This time you will obtain black lacquer, which should cover quite well and be very durable.

Water paints have a wide field of usefulness, principally of course, for indoor work where they will not be subject to the action of the weather. An exception is whitewash which, even though mixed with water, will stand up outdoors for quite a long time.

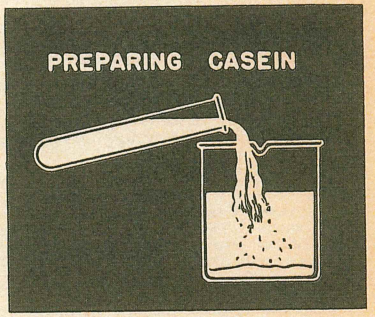
EXPERIMENT 308—Making Whitewash.

Put 5 measures of Calcium Oxide (No. 20) in a cup and add 5 or 6 drops of water. Stir this mixture until it forms a smooth thin paste. This is whitewash. Whitewash is used extensively for painting fences and rough wooden buildings. The particles of lime fill the pores of the wood and since the lime is not soluble in water, the coating will last for quite a long period of time. If desired, a little glue can be added to the whitewash to make it stick better when applied to a surface.

Paints which will stand up under fairly severe conditions can be made, using a solution of casein in water together with any desired pigment to add color. Freshly prepared casein is readily taken up by water forming a colloidal suspension or solution, especially if a little lime or other alkaline substance is present. After drying, however, the casein becomes insoluble, so that a wall painted with casein paint can be washed after the paint has thoroughly dried.

EXPERIMENT 309—Preparing Casein.

Fill a tumbler about ¼ full of skim milk. In a test tube half full of water dissolve 10 measures of Sodium Bisulfate (No. 7). Add this solution, a few drops at a time, to the skim milk, stirring the milk continually with a spoon. After the sodium bisulfate has all been added to the milk, you will notice a gelatin-like precipitate being formed in the milk and upon allowing it to stand for a while, the precipitate becomes heavy. This precipitate consists principally of casein. Carefully pour off the liquid, saving the casein in your glass for a future experiment.



EXPERIMENT 310—Making Brown Casein Paint.

Prepare some brown iron oxide pigment as in a previous experiment and carefully mix this pigment together with ½ measure of Calcium Oxide (No. 20) and some of the casein which you have just prepared. You will find that the casein is very soft and that the brown pigment can be ground into it, making a smooth paint. This paint can be thinned with water, if desired. Once your casein has dried, it cannot be again mixed with water as it has become insoluble.

EXPERIMENT 311—Making Black Casein Paint.

Proceed as in the above experiment, but use lampblack as a pigment instead of brown ferric oxide.

EXPERIMENT 312—Making Light Blue Casein Paint.

Mix 5 or 6 measures of Azurite (No. 39) and ½ measure of Calcium Oxide (No. 20) with a small amount of your soft casein. This will make a very pretty blue paint, which can be thinned to the desired consistency with water.

Poster colors consist of pigments ground with a clear mucilage or glue. These colors are quite brilliant and in practice, they are usually applied to white or colored cardboard. They are sometimes called show-card colors.

EXPERIMENT 313—Purple Lake Poster Color.

In this and the following experiments you will need a small bottle of clear mucilage with which to mix your colors. Clear mucilage can be obtained in almost any five and ten cent store or drugstore.

Put 2 measures of Logwood (No. 29) in a test tube half full of water and heat the tube for 3 or 4 minutes to extract the color. When the water in the tube has become a deep brown color, add 1 measure of Aluminum Sulfate (No. 12). Shake the tube thoroughly and carefully pour the clear red solution into another test tube, leaving behind the chips of logwood and any other sediment. Add to this clear red solution 1 measure of Sodium Carbonate (No. 4) and again shake the tube. The result will be a deep purple precipitate. Let this tube stand overnight until the precipitate settles, then pour off the clear liquid and scrape the purple precipitate into a small cup or onto a flat piece of glass. This purple pigment, when mixed with clear mucilage, will make a very pretty purple poster paint.

In this and several following experiments much time can be saved by filtering the liquid to separate the precipitate. See the instructions for filtering under "The Use of Chemical Apparatus."

EXPERIMENT 314—Making Black Lake Poster Color.

A black pigment, very similar to the purple pigment which you just made, can also be prepared from logwood. Heat 2 measures of logwood in a test tube half full of water for several minutes until you extract the color from the logwood. Add 1 measure of Ferric Ammonium Sulfate (No. 21) and shake the tube thoroughly. Pour off the black solution which is formed, being careful to leave behind the chips of logwood and any other sediment. Now add to this black solution, 1 measure of Sodium Carbonate (No. 4). Shake thoroughly and then let the black precipitate settle, so that you can pour off the clear liquid on top. This black pigment, when mixed with clear mucilage, makes an excellent black poster color.

EXPERIMENT 315—Light Blue Poster Color.

Mix 3 or 4 measures of Azurite (No. 39) with a few drops of mucilage. The result will be a pretty light blue poster color.

EXPERIMENT 316—Brown Poster Color.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. Dissolve 4 measures of Sodium Carbonate (No. 4) in another test tube ⅓ full of water. Pour the first solution into the second and you will obtain a thick brown precipitate which consists of ferric hydroxide. Let this precipitate settle, pour off the clear liquid from the top, and then scrape the precipitate from the test tube into a cup or onto a glass plate. This brown pigment, when mixed with clear mucilage makes a rich brown poster color.

EXPERIMENT 317—Green Poster Color.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ⅓ full of water. Now add 2 or 3 drops of Cobalt Chloride Solution (No. 27). Shake the tube and you will obtain a thick green precipitate of cobalt ferrocyanide. Allow this precipitate to settle, pour off the clear liquid on the top and scrape the green pigment onto a glass plate and mix with clear mucilage to form green poster color.

EXPERIMENT 318—Red-Brown Poster Color.

Put 2 measures of Azurite (No. 39) and 4 measures of Sodium Bisulfate (No. 7) in a test tube and fill it half full of water. Warm the tube for a few moments until the solution becomes a clear blue color and all the solids dissolve. In another test tube, dissolve 2 measures of Sodium Ferrocyanide (No. 6), shaking to dissolve all of the solid. Now pour the contents of the first tube into the second and you will obtain a red-brown precipitate of copper ferrocyanide. By allowing this precipitate to settle, pouring off the clear liquid and then scraping the precipitate onto a glass plate and mixing it with clear mucilage, you will obtain a reddish-brown poster color.

EXPERIMENT 319—Lampblack Poster Paint.

Proceed as in the previous experiments but use lampblack as a pigment and mix thoroughly with the clear mucilage. This will give a dead black poster paint.

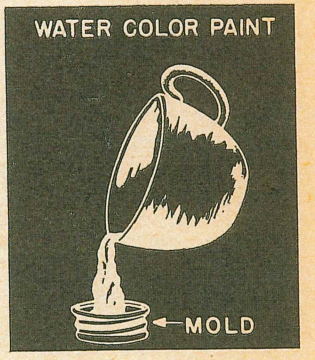
EXPERIMENT 320—White Poster Paint.

Zinc oxide makes one of the best white pigments for poster paint. Precipitated chalk or chalk dust can also be used as a pigment for white poster paint.

Artists' water color paints usually consist of a pigment or dye mixed with glucose, or a similar water soluble base. The glucose, or other base, is often heated while the color is added and upon cooling, the base then solidifies into a solid cake. Ordinary corn syrup is glucose and if you have some of this available, you can easily make a few water color paints.

EXPERIMENT 321—Blue Water Color Paint.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. In another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). Pour the second solution into the first and a deep blue color will result. Put a tablespoon of corn syrup into an old china cup and float the cup on a pan of hot water until the corn syrup becomes warm and fairly fluid. Now add a little of the blue color which you have prepared and mix it thoroughly with the warm corn syrup. This mixture can be poured into small molds and after it stands for a few hours, the cakes of paint will be hard just like regular water colors.



EXPERIMENT 322—Another Blue Water Color Paint.

Cut a small piece of Dye Wafer (No. 8) about ¼ inch wide by ½ inch long. Put this in a test tube ¼ full of water and shake the tube for a few minutes to extract the dye. Now warm some corn syrup as in the preceding experiment and mix the blue liquid from the dye wafer with the glucose. Pour it into molds and allow it to solidify as before.

EXPERIMENT 323—Black Water Color Paint.

Dissolve 1 measure of Tannic Acid (No. 15) in a test tube ¼ full of water. In another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). When the contents of both test tubes are dissolved, pour the tannic acid solution into the ferric ammonium sulfate solution and you will obtain an intense black color. Mix some of this black color with warm corn syrup and pour this mixture into molds to solidify. You will have cakes of black water color paint.

EXPERIMENT 324—Yellow Water Color Paint.

Put 2 measures of Logwood (No. 29) in a test tube ⅓ full of water and heat the tube for 3 or 4 minutes until the water has become colored by the logwood. Add ½ measure of Sodium Bisulfate (No. 7). Now add some of this yellow dye to warm corn syrup as before and the result will be a yellow water color paint.

EXPERIMENT 325—Other Colors of Water Color Paint.

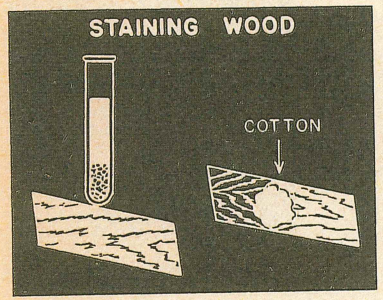
There is no limit to the colors that you can make up in these paints if you have water soluble dyes of various colors and shades. Perhaps you may have some packages of Easter egg dyes or dyes used in coloring fabrics; even colored inks can be used. Follow the directions in the preceding experiments in making up various other colors of water color paints.

Stains consist of soluble dyes or coloring material dissolved in a liquid. Stains for use out of doors where they are exposed to the weather are usually made with naphtha, turpentine, lacquer thinner or other organic solvent. Stains for use indoors can be made with water, although for woodwork, floors and other places where washing is necessary, the organic solvents are usually best.

Stains cannot be used on surfaces into which the liquid cannot penetrate. Thus, the principal use of stains is on wood. Woodwork is often stained to give the desired color or tone and then covered with a clear varnish or lacquer, which gives a glossy finish.

EXPERIMENT 326—Brown Lacquer Thinner Stain.

If you can obtain some nail polish remover, fill a test tube about ¼ full and add 2 measures of Logwood (No. 29). Use a cork, or crumple up a small piece of paper to stopper the mouth of the test tube and set it aside for several hours. The lacquer thinner will gradually take up color from the logwood forming, after some time, a very pretty yellowish brown stain that can be applied to any white wood with good effect.



EXPERIMENT 327—Brown Alcohol Stain.

Fill a test tube ¼ full of alcohol (rubbing alcohol will be satisfactory). Add 2 measures of Logwood (No. 29) and let this soak in the alcohol for an hour or so. This will form a brown stain which penetrates evenly into the pores of wood and makes a very attractive finish.

EXPERIMENT 328—Blue Lacquer Thinner Stain.

Cut a small piece of Dye Wafer (No. 8) about ¼" wide by ½" long and place this in a test tube ¼ full of nail polish remover. The lacquer thinner will gradually dissolve the blue dye and if left standing for some time, will produce a blue stain.

EXPERIMENT 329—Blue Alcohol Stain.

A small piece of Dye Wafer (No. 8) can be used with alcohol to produce a blue stain. The alcohol will dissolve the color better than the lacquer thinner and will make a more intense stain.

EXPERIMENT 330—Blue Water Stain.

Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. In another test tube ⅓ full of water, dissolve 1 measure of Sodium Ferrocyanide (No. 6). Pour the second solution into the first and you will obtain a blue color which can be used for staining wood or other absorbent material.

EXPERIMENT 331—Black Water Stain.

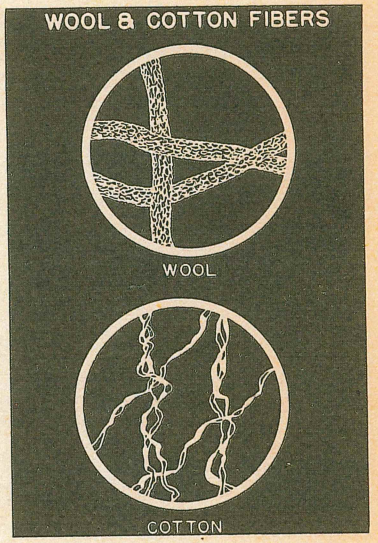
Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Dissolve 1 measure of Tannic Acid (No. 15) in another test tube ¼ full of water. When these substances are all dissolved, pour the tannic acid solution into the ferric ammonium sulfate solution and you will obtain a black color which can be used as a water stain.

EXPERIMENT 332—Yellow Logwood Stain.

Put 2 measures of Logwood (No. 29) in a test tube ¼ full of water. Warm the liquid for 3 or 4 minutes and add ½ measure of Sodium Bisulfate (No. 7). This makes an excellent stain for use on a light colored wood.

TEXTILE FIBERS

The textile industry is of great importance in our modern world. The business of growing or making the fibers, spinning them into thread, weaving the thread into cloth and making the cloth into useful articles employs a great many people. The output of the textile industry touches our daily life in many ways. Our clothing, rugs, bedding, curtains, the cords in automobile tires, the insulation of electric wires, cords and ropes of all kinds from cotton string to the great hawsers of ships, the binding of books, and dozens of other indispensable articles are products of the textile industry.



There are three general classes of textile fibers:

1—Fibers obtained from animals, such as wool, silk and fur.

2—Fibers obtained from plants, such as cotton, flax, jute, hemp and manila.

3—Fibers made from synthetic products, such as rayon, estron, nylon, casein and glass.

NATURAL FIBERS

Vegetable fibers, such as cotton and linen, consist chiefly of cellulose and are decomposed by strong acids. Animal fibers, such as silk and wool, consist largely of compounds containing nitrogen, carbon, hydrogen and sometimes sulfur. These withstand the action of acids but are readily destroyed by alkalies.

Many people can tell whether a fiber is silk, wool or cotton by feeling it and examining it closely. It is impossible, however, to detect certain mixtures of fibers in cloth by these methods, and so standard chemical tests have been devised by means of which the exact composition of cloth can be determined. Examination under the microscope is one of the best ways to distinguish these fibers.

EXPERIMENT 333—Testing Fabrics for Animal Fibers.

Animal fibers, such as silk and wool, burn slowly, giving an odor similar to burnt hair or horn, and form round granules of carbon at the point of the thread where combustion ceases.

EXPERIMENT 334—Testing Fabrics for Vegetable Fibers.

Vegetable fibers, such as cotton, artificial silk, and linen, burn rapidly, form a thin, white ash and give off only a faint smell, similar to burnt paper.

Take threads from various kinds of fabrics and burn them, watching the results closely. This is the simplest test for determining various fabrics.

EXPERIMENT 335—Testing Fabrics for Pure Linen Fibers.

Dissolve ½ measure of Sodium Bisulfate (No. 7) in a test tube half full of water. Put a small scrap of linen in this solution and boil it for a few minutes. Remove the linen, wash it carefully in water and then dry it.

Now fringe out the edges of the piece of cloth so as to expose the fibers and moisten one corner with a drop of glycerine. Press the cloth between blotters and then lay it against a dark background. The linen fiber will be translucent and if any other kind of fibers are mixed with the linen, they will appear opaque.

EXPERIMENT 336—The Alkali Test for Animal Fibers.

Wool and natural silk are dissolved by boiling them in sodium hydroxide solution while cotton, linen and artificial silk are not affected.

Prepare a solution of sodium hydroxide by dissolving 2 measures of Calcium Oxide (No. 20) and 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water. Boil this solution for a few minutes, and then put in a small scrap of wool or silk and boil for a few minutes more. Be careful that the liquid does not spurt out of the tube into your face. Notice the effect on the fibers. In case a fabric is a mixture of wool and cotton, the wool fibers will be dissolved and the cotton fibers will remain.

SYNTHETIC FIBERS

At present the most important synthetic textile fibers are rayon, estron and nylon. These fibers have almost replaced natural silk and they are undoubtedly making inroads on the other natural fibers.

The basic material for both rayon and estron is cellulose. The cellulose may be in the form of specially purified wood pulp, or of cotton linters, that is, the short fibers removed from the cotton seeds by a second ginning.

Most of our rayon is made by the so-called viscose method, in which the purified cellulose is treated with sodium hydroxide solution and with carbon disulfide. The reaction which occurs forms a yellow spongy mass of cellulose xanthate. When more sodium hydroxide is added, this becomes a syrupy liquid, which is forced through tiny holes in a metal plate into a bath of sulfuric acid. When the filament of liquid comes in contact with the sulfuric acid, the sodium hydroxide in it is neutralized and the cellulose xanthate is decomposed, giving back the original cellulose, but in the form of fine, uniform fibers of unlimited length. These may be twisted into thread.

Two less-used processes, the cellulose nitrate and the cuprammonium, differ only in the means used to get the cellulose into liquid form. In both of them the liquid is converted into cellulose at the end of the process.

Cellulose acetate, formerly called acetate rayon, now known as estron, differs from that made by the viscose process in that it is not cellulose. The starting material is cellulose, just as in the other process. Treatment with strong acetic acid converts the cellulose into cellulose acetate. This is dissolved in acetone, squirted through fine nozzles to form it into fibers and the acetone removed. It remains, however, in the form of cellulose acetate instead of being converted back into cellulose.

Nylon was first produced commercially in the late 1930's. In chemical composition, it differs greatly from rayon and estron since it contains nitrogen in addition to carbon, hydrogen and oxygen. Chemically, it is more like natural silk. By a complicated series of reactions, suitable compounds are obtained from basic raw materials — coal, air and water. These compounds are melted together and slowly react with one another, forming long chains of atoms. The melted material is forced through tiny holes just as in making rayon and it then hardens in the form of filaments which, after a stretching process, are spun into thread.

Other synthetic fibers include polyvinylidene chloride under the trade name of saran, casein fibers made by extruding casein into a solution of formaldehyde, and polyacrylonitrile fiber, which is being produced under the trade name of orlon. In the section on glass, we mentioned the very fine glass fibers which it is now possible to make. They are finding increasing application in the textile field.

The various synthetic fibers have quite different physical characteristics, which makes them adaptable for different purposes.

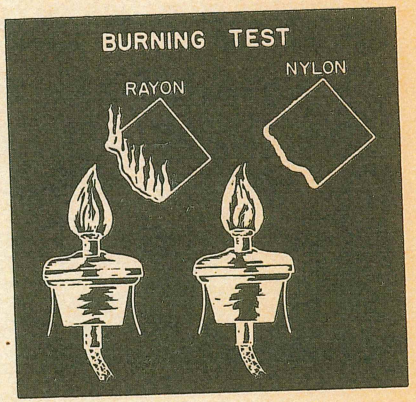
Viscose rayon is very susceptible to moisture changes and material made from these fibers will swell and absorb up to 75% of its weight of water. On the other hand, viscose rayon has high tensile strength and is quite resistant to heat, which makes it ideal for use in automobile tire cord and for many other purposes.

Cellulose acetate, estron, has a considerably lower moisture absorption and is faster drying. It is, however, much more sensitive to heat, melting at a considerably lower temperature.

Nylon combines high tensile strength, low moisture absorption and a relatively high melting point, which makes it suitable for purposes where great strength and resistance are required.

EXPERIMENT 337—The Burning Test for Synthetic Fibers.

Cut a small square of rayon or estron and of nylon from worn-out clothing or rags. Hold the edge of each square in a flame for a moment. If your sample is nylon, it will cease to burn as soon as removed from the flame. If it is either rayon or estron, it will continue to burn slowly when removed from the flame. If it is estron (cellulose acetate), the smoke which comes from the flame will have a characteristic sharp odor. If it is rayon (viscose), it will continue to burn slowly, the smoke smelling somewhat like burning paper.



EXPERIMENT 338—Melting Point of Synthetic Fibers.

Test small samples or rayon, estron and nylon under a hot iron. The cellulose acetate, or estron, will melt first and the viscose and nylon will stand a considerably higher temperature.

EXPERIMENT 339—Solubility of Synthetic Fibers.

Rayon, estron and nylon are not soluble in ordinary petroleum derivatives, such as naphtha or gasoline. They are all soluble, however, in other organic solvents such as lacquer thinner. This may be demonstrated by placing small squares of cloth made from different synthetic fibers in a test tube and adding a small amount of nail polish remover. Thus, care must be taken not to moisten cloth made of these fibers with liquids such as lacquer thinners.

In many ways the synthetic fibers are superior to our natural fibers, but on the other hand, natural fibers have certain important physical characteristics which so far we have been unable to duplicate in the synthetic fibers.

DYES

In the process of dyeing, the coloring matter is dissolved in water or other solvents and the material to be dyed passed through the solution.

Many dyes unite directly with the fiber of the cloth, forming, in some cases, colors which are insoluble and will withstand washing; in other cases, the colors formed in this manner are not entirely "fast" and will "run" to some extent when the cloth is washed.

Since there are many different kinds of fibers used in making textiles, we are not surprised that the same dye gives different shades and intensity

of color on cloth made from different fibers. A dye which takes well on wool, for example, might be entirely unsuited for saran. Sometimes when two or more fibers are used in weaving the same piece of cloth, certain dyes will give a variegated effect.

There is also a great variety of coloring matters which are dyed on fabrics by the aid of metallic mordants. A mordant is a substance which will combine both with the fiber of the cloth and with the dyestuff. Salts of aluminum, iron and tin are often used. Tannin, or tannic acid, is another valuable mordant. The cloth to be dyed is first passed through a solution of the mordant which unites chemically with the fiber, then through the dye solution, the dye in turn uniting with the mordant and forming a "fast" color. The same dyestuff frequently gives different shades of color when used with various mordants.

EXPERIMENT 340—Dyeing Fabrics Dark Blue.

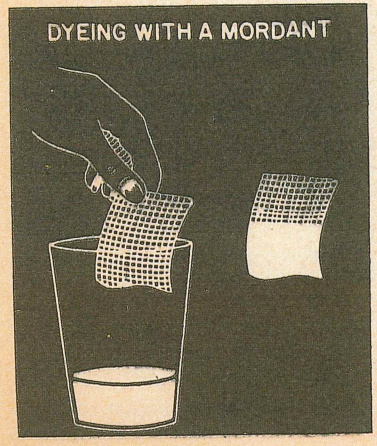
Prepare a solution of 4 measures of Sodium Ferrocyanide (No. 6) in a test tube half full of water. Cut a small piece of absorbent cloth about 1 inch or ½ inch square and place it on the bottom of a clean tumbler. Now pour the solution of sodium ferrocyanide into the tumbler and saturate the cloth thoroughly with the solution. Remove the cloth and while it is drying, prepare the following solution: 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. When the cloth has dried, place it in the bottom of a clean tumbler and pour this solution over it. Stir it around well so that the solution wets every part of it. Now remove the cloth from this solution and you will find that it is dyed a dark blue color.

EXPERIMENT 341—Dyeing Fabrics Black.

Prepare a solution of 2 measures of Tannic Acid (No. 15) in a test tube half full of water and moisten a small square of cloth with the solution as in the preceding experiment. When the cloth becomes thoroughly dry, moisten it again with a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. This time the cloth will be dyed black.

EXPERIMENT 342—Dyeing Fabrics Brown.

Prepare a solution of 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water and moisten a small square of cloth with the solution as in the preceding experiment. When the cloth becomes thoroughly dry, moisten it again with a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. The cloth will be dyed brown.



EXPERIMENT 343—Dyeing with a Mordant.

Put 4 measures of Logwood (No. 29) in a glass and pour in one test tube full of hot water. Let the solution stand half an hour to extract the color.

Cut a small slip of white cloth about 2 inches square and wet one corner of it with a solution of 2 measures of Aluminum Sulfate (No. 12) in half a test tube of water. Put this where it will dry.

After the logwood solution has stood for half an hour, add 3 measures of Sodium Carbonate (No. 4) to it and stir well. Now put the piece of cloth, which must be dry, into this solution and let it soak for about 10 minutes.

When the cloth has been taken out and dried, you will see that it has been dyed two colors, the part which was treated with a solution of aluminum sulfate being darker than the rest of the cloth.

EXPERIMENT 344—Dyeing Cotton Cloth Buff.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. Put a small square of cotton cloth in this solution and allow it to soak for 5 or 10 minutes. Remove the cloth, wring it until nearly dry and then put it in a solution of 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water. Allow it to soak in this solution for about 5 minutes. Again wring the cloth nearly dry, and then soak it in a solution made by boiling a small piece of soap in a test tube half full of water. The cloth will be dyed buff, and a deeper shade can be obtained by repeating the operation with the same piece of cloth.

A series of organic coloring materials may be prepared by heating various organic substances with sulfur in the presence of an alkali. These sulfur dyes, as they are called, are frequently used in dyeing cotton fabrics.

EXPERIMENT 345—The Manufacture of Sulfur Black.

In a dry test tube put 1 measure of Tannic Acid (No. 15), 1 measure of Sodium Carbonate (No. 4) and 1 measure of Sulfur (No. 1). Shake the tube gently to thoroughly mix the powder and heat over a flame for 3 or 4 minutes. Keep your face away from the mouth of the test tube. Set the tube aside to cool.

When the tube is cool, fill it ¾ full of water, close the mouth with your thumb and shake well. Pour the dark solution obtained into a glass, add fresh water to the tube and shake again to dissolve any remaining dye. Notice the intense coloring power.

EXPERIMENT 346—Making Sulfur Brown.

Put 1 measure of Sulfur (No. 1), 1 measure of Sodium Carbonate (No. 4) and 3 or 4 drops of glycerine in a test tube and heat over the alcohol lamp flame for 3 or 4 minutes. Keep your face away from the mouth of the tube.

When the test tube has cooled, shake the dye with a little water to loosen it from the test tube, and pour the dye solution into a tumbler. When you add more water to this dye solution you will find that it is brown.

EXPERIMENT 347—Making Green Sulfur Dye.

Place in a test tube 2 measures of Sulfur (No. 1) and 2 measures of Sodium Carbonate (No. 4). Add about 1 measure of butter, and heat over a flame. At first the sulfur and butter will melt, then the mixture will turn dark and become foamy. When the contents of the test tube have reached a good dark color, stop heating and allow to cool. Then fill the tube about half full of water and shake vigorously. The precise shade of green which you can obtain will depend upon the amount of butter you use and the length of time you heat the mixture.

Natural dye materials, such as logwood, were at one time very largely used in the dyeing industry, but at the present time, the use of the aniline dyes and other artificial dyestuffs has largely replaced the naturally occurring dye materials.

EXPERIMENT 348—The Effect of Acids and Bases on Logwood.

Place 2 measures of Logwood (No. 29) in a test tube and fill the tube ½ full of water. Warm a few minutes to extract the color. In another test tube ¼ full of water dissolve 1 measure of Sodium Bisulfate (No. 7).

Pour the logwood solution into the sodium bisulfate solution and a yellow color results.

Now add 2 measures of Sodium Carbonate (No. 4). The solution effervesces and gradually changes to a dark purple color.

EXPERIMENT 349—Logwood Black.

Warm 2 measures of Logwood (No. 29) in a test tube ½ full of water as in the preceding experiment. In another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21).

Pour the logwood solution into the ferric ammonium sulfate solution and a very intense color results.

EXPERIMENT 350—Dark Red Logwood Color.

Warm 2 measures of Logwood (No. 29) in a test tube ½ full of water as before and add to this 3 or 4 drops of Cobalt Chloride Solution (No. 27) and ¼ measure of Sodium Carbonate (No. 4). This time a very pretty dark red color is formed.

EXPERIMENT 351—Olive Green Logwood Color.

Warm 2 measures of Logwood (No. 29) in a test tube ½ full of water as before. In another test tube ¼ full of water, dissolve 1 measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7). Pour the logwood solution little by little into this test tube and notice the olive green color which gradually appears.

EXPERIMENT 352—Buckeye Bark Dye.

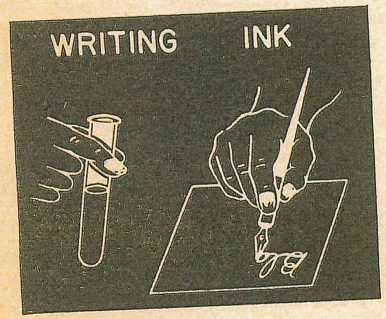
Place a few pieces of the bark of the horse-chestnut or buckeye tree in a test tube filled with water to which a small quantity of household ammonia has been added. In a few minutes streamers of blue fluorescent dye will descend from the chips of bark.

EXPERIMENT 353—Osage Orange Wood Dye.

The wood of the osage orange tree contains a valuable yellow dye. If you can obtain a few chips of this wood, boil them with a little water to extract the color.

THE MANUFACTURE OF INKS

The manufacture of writing ink is an age-old industry, and the first inks consisted of fine soot mixed with a solution of some kind of gum. Today there are many kinds of ink for every conceivable purpose, and many formulas are in existence for the manufacture of each kind.



EXPERIMENT 354—How to Make Writing Ink.

Fill a test tube ⅓ full of water and add ½ measure of Tannic Acid (No. 15). Put an equal amount of water in another test tube and add ½ measure of Ferric (Iron) Ammonium Sulfate (No. 21). Shake both test tubes to dissolve the solids and then pour the contents of the first into the second.

A deep black color will appear, due to a very finely divided tannate of iron which was precipitated when the liquids were mixed and which remains in suspension.

The manufacture of most of our writing inks is based on this reaction, though many of them contain, in addition, a certain amount of some aniline dye. Gum arabic is also generally added to thicken the ink slightly and prevent the suspended iron tannate from settling out.

These iron gall inks, as they are called, have the advantage of being permanent, that is, they withstand the action of air and light for many years, and they are very cheap to manufacture. They are, however, very easily erased by the use of certain chemicals, and steel pens are more or less corroded by them.

EXPERIMENT 355—How to Make Blue Ink.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. In another test tube ⅓ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). Pour the second solution into the first. A deep blue liquid will result, due to the formation of Prussian blue. This is widely used as a blue ink. Try writing with it. Prussian blue is often used as a substitute for indigo in laundry bluing.

EXPERIMENT 356—How to Make Violet Ink.

Put 2 measures of Logwood (No. 29) in a test tube and fill the tube ¼ full of water. Warm for 3 or 4 minutes to extract the color and then add ½ measure of Aluminum Sulfate (No. 12) to the hot solution. When the aluminum sulfate has dissolved, write with the solution, using a clean pen.

EXPERIMENT 357—How to Make Red Ink.

Put 2 measures of Logwood (No. 29) in a test tube and fill the tube ¼ full of water. Warm 3 or 4 minutes to extract the color. Then add ½ measure of

Aluminum Sulfate (No. 12) and 1 measure of Sodium Bisulfate (No. 7). When these solids have completely dissolved, try writing with the ink, using a clean pen.

EXPERIMENT 358—How to Make Green Ink.

To make some green ink place about 3 or 4 measures of shaved barberry hedge or osage orange wood in a test tube and add about ½ inch of water. You can obtain the barberry wood or osage orange from your hedge or you may ask around your neighborhood for a small trimming from either of these bushes.

Next warm the test tube containing the wood and water to extract the yellow color. After a minute of boiling, add 1 measure of Sodium Ferrocyanide (No. 6) and a tiny crystal of Ferric Ammonium Sulfate (No. 21).

This will give you a deep green ink. Try writing with it, using a clean steel pen.

EXPERIMENT 359—Making Brown Ink.

In a test tube half full of water, put 1 measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7). Shake this solution for a moment until both solids are dissolved, then, add 1 measure of Sodium Ferrocyanide (No. 6) and shake the solution again. This gives a pretty brown colored ink.

EXPERIMENT 360—Preparation of Permanent Ink.

Fill a test tube ⅓ full of water and add 1 measure of Tannic Acid (No. 15) and 1 measure of Ferric Ammonium Sulfate (No. 21). Shake this solution thoroughly.

In another test tube ⅓ full of water, put 2 measures of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21). Shake this solution thoroughly.

Now pour the second solution into the first one. Shake the test tube thoroughly and try the ink which you have prepared. If you wish to keep this ink, add a drop of carbolic acid to prevent it from molding.

EXPERIMENT 361—Making a Black Ink Powder.

Mix together 1 measure of Tannic Acid (No. 15) and 1 measure of Ferric Ammonium Sulfate (No. 21). Place the powder thus formed in a small bottle; and when you need ink, simply add some of this powder to a little water.

EXPERIMENT 362—Making a Blue Ink Powder.

Blue ink powder may be made by mixing together 2 measures of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21). When this powder is mixed with water, you have blue ink for writing.

Chemists have long endeavored to discover an ink which cannot be erased by any known means. The value of such an ink is self-evident. Documents written with it would not only resist the action of time for thousands of years, but would also be absolute proof against change. Unfortunately, there is no known formula which will yield a satisfactory ink which cannot be erased by some means. Printers' ink is the only exception, as in this case finely divided carbon is so firmly pressed into the pores of the paper that it cannot be removed, except by seriously damaging the paper. A safety ink which is fairly good may be made as follows.

EXPERIMENT 363—Making a Safety Ink.

Place a spoonful of Sodium Silicate Solution (No. 19) in a dish and stir in 1 measure of lampblack. This mixture must be stirred for some time with a glass rod in order to properly mix the ingredients.

You will find that it will be rather difficult to write with this ink as it does not flow very well, and it must be bottled up in a tightly stoppered bottle if it is desired to keep it any length of time. Writing done with this ink, however, is very difficult to remove, as the lampblack is held firmly by the sodium silicate solution.

EXPERIMENT 364—Another Type of Safety Ink.

Dissolve 1 measure of Tannic Acid (No. 15) and 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. Shake this solution thoroughly and then add 1 spoonful of Sodium Silicate Solution (No. 19). Pour this mixture

into a cup and add 2 measures of lampblack. Now mix this thoroughly with a spoon. The result will be a deep brown or black ink which is very durable and which is almost impossible to erase without damaging the surface of the paper.

EXPERIMENT 365—How to Make an Ink Remover.

Put 1 measure of Sodium Bisulfate (No. 7) and ½ measure of Bleaching Powder (No. 38) in a test tube and add ½ inch of water. Shake the tube slightly to dissolve the solids.

Now place a few drops of this solution on some black writing. Spread it around with a soft brush or piece of cotton. If the writing was done with ordinary ink, it will soon be entirely bleached so that you could not possibly tell that the paper had been written on. If, however, the writing was done with an ink containing Prussian blue, it will not be erased with this ink remover.

EXPERIMENT 366—How to Make a Substitute for Carbon Paper.

Obtain a cake of white soap and by scraping the soap with a knife, make a pile of fine soap shavings. Put 15 measures of the soap shavings in a glass or cup. In a test tube ⅓ full of water, put 1 measure of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21), and shake thoroughly until the solids are all dissolved and you have a blue solution. Pour this blue solution into the glass or cup with the soap shavings and stir thoroughly for the next five minutes until the mixture forms a smooth, thick paste.

Now using a small brush or piece of absorbent cotton, spread this paste smoothly over a piece of paper. Set the paper away to dry for at least a day and then use it as you would use ordinary carbon paper.

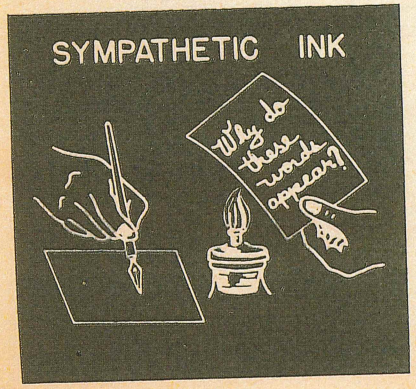
EXPERIMENT 367—How to Make a Blue Typewriter Ribbon.

Obtain a quantity of soap shavings by scraping a piece of white soap with a knife. Put 20 measures of these soap shavings in a cup and add 4 or 5 drops of glycerine.

Now dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) and 1 measure of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. Shake this solution well so that the solids will be entirely dissolved and then pour it into the cup containing the soap shavings and glycerine. Stir this mixture until it forms a smooth paste.

Get a piece of ribbon or tape about the width and thickness of a typewriter ribbon and apply the paste which you have made evenly to the cloth or tape. Now hang the tape up to dry for half an hour or so and then roll it into a roll just like a typewriter ribbon. In order to use it on a typewriter you will, of course, have to roll it on an empty spool of the same kind used on your typewriter.

In time of war important messages are often written in secret inks. An apparently innocent letter may contain vital information written between the lines with invisible ink and this message can be made available only to those who possess the secret of developing the ink.



EXPERIMENT 368—Making a Sympathetic Ink.

Using a clean pen, write on white paper with Cobalt Chloride Solution (No. 27). This writing will be invisible until the paper is warmed when it will appear in blue characters. When the paper is again allowed to cool, the writing will disappear.

EXPERIMENT 369—How to Make a Permanent Sympathetic Ink.

Dissolve 4 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Using a clean pen, write with this solution on white paper. The writing cannot be seen until the paper is heated, when it appears in dark brown lines. This writing is permanent and will not disappear.

EXPERIMENT 370—Lemon Juice Ink.

Try writing with the juice of a lemon. This writing will be invisible until the paper is warmed, when it will appear permanently in brown lines.

EXPERIMENT 371—Another Type of Sympathetic Ink.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. Write with this solution, using a clean pen. If you are careful, the writing will be invisible. To develop this writing, dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. Moisten a piece of cotton or a soft brush with this solution and carefully paint over the invisible writing on the piece of paper. The writing will appear permanently in blue lines.

Fingerprints today are a very important means of identification in crime detection. The Federal Bureau of Investigation, as well as many metropolitan police departments, have very large files of fingerprints which are of great value in identification work and in criminal investigation. You can make your own fingerprint ink and obtain fingerprints of your friends.

EXPERIMENT 372—How to Make Fingerprint Ink.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. In a second test tube ¼ full of water, dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). Pour the second solution into the first and a deep blue solution will result due to the formation of Prussian blue.

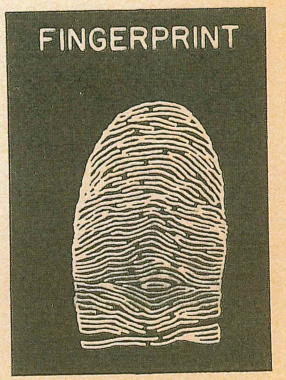
With this colored solution you can make very clear fingerprints, the method being explained in the following experiment.

EXPERIMENT 373—How to Make Fingerprints.

Saturate a small piece of absorbent cotton or soft cloth with the colored solution made in the preceding experiment. Lightly paint the ball of your left index finger with the cotton, wipe off the excess, and then press it against a sheet of white paper. Lift your finger from the paper and you will find a fingerprint different from every other one in the world. No two fingerprints are entirely identical.

Often a better fingerprint will result if the finger is slowly rolled on the ball from the left to the right side and then lifted clear of the paper.

To obtain a complete set of your fingerprints do the remainder of your fingers and thumb on your left hand. After this is finished, make fingerprints of the thumb and fingers of your right hand. Always wash your hands with soap and water after making fingerprints.



EXPERIMENT 374—Making a Fingerprint Album.

An interesting album can be compiled by making complete sets of fingerprints of all your friends.

On the upper half of a sheet of white paper make the complete set of "prints" from the right hand. Using your blue fingerprint ink and a clean steel pen, write or print "RIGHT HAND" beneath the fingerprints just made.

Immediately beneath this, make the "prints" from the fingers and thumb of the left hand and title them "LEFT HAND."

On the lower part of the sheet you should list the person's name, the date and any other data which you may wish for future reference.

After making the fingerprints of your friends, always have them wash their hands with soap and water.

Telltale fingerprints which are sought for so diligently by detectives and crime investigators are usually invisible unless treated with some form of

fingerprint powder which causes them to become visible. Those found on light surfaces are made visible with a dark powder.

EXPERIMENT 375—Detecting Fingerprints on Light Surfaces.

Obtain a smooth piece of paper, breathe on one of your fingers a few times and press it on the paper. The print of your finger is not visible. Now prepare a mixture of 3 parts of lampblack and 3 parts of Sulfur (No. 1). Put this mixture on the paper which you have touched with your finger and tap the underneath side of the paper gently, so that the powder will be spread over the print. Shake off the excess powder and you will find that the fingerprint which was invisible is now plainly seen. To "set" the print, warm the paper gently over a flame, being careful not to burn it.

This is the way "G-Men" often develop old or latent fingerprints.

EXPERIMENT 376—Detecting Fingerprints on Dark Surfaces.

Grasp a dark bottle in your hand. Now take 2 measures of Sulfur (No. 1) and 4 measures of white talcum powder, and mix them thoroughly.

Using the method described in the preceding experiment, shake some of the powder over your invisible fingerprints. Immediately the prints become visible. Blow off the excess powder and warm the bottle gently to set the prints. After a little experience, you will be able to bring out fingerprints on pens, chinaware, windows and many other objects.

ADHESIVES

Adhesives are materials which, when applied to the surfaces of substances, cause them to adhere to one another. There are thousands of formulas for various kinds of adhesives, and special types have been developed for practically every industry requiring a material of this kind.

Animal glue is made from clippings from the hides, horns and hoofs of animals. These are soaked in acid or alkali to soften them and then heated in water to set free the gelatin. Animal glue is commonly sold in the form of flakes, sheets or powder for use in making furniture, veneered wood, paper boxes, carpets, coated papers, calcimine, and many other materials. Fish glue is made from the wastes of the fish industries. It is generally sold in the liquid form for the use of bookbinders and others requiring a flexible adhesive.

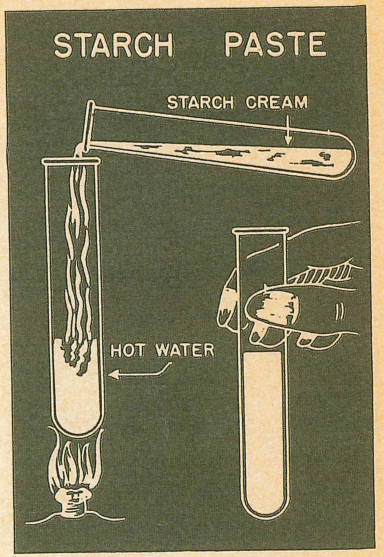
Casein glue is manufactured from crude casein obtained from skim milk under very careful chemical control. This glue is superior to animal glue in its water resisting qualities. Using casein as a base, it is possible to make a glue that is soluble in cold water, but becomes waterproof when hardened.

Another type of adhesive contains cellulose acetate dissolved in a suitable solvent. Often this solvent is amyl acetate, which will be recognized by its odor similar to that of bananas. Because of the volatility of the solvent, this type of adhesive is quick drying and frequently used in the construction of model airplanes. Many other kinds of plastic adhesives that have come into general use include phenol formaldehydes and urea formaldehydes which are widely used in industry for special purposes.

An adhesive known as "rubber cement" is made by dissolving crude rubber in a solvent. These solvents are usually either benzine or carbon tetrachloride, both being very volatile. Rubber cement is ideal for attaching paper to paper, to glass, and to wood. This cement does not cause paper to buckle as is customary with mucilages, and because of this property, finds wide usage among artists, photo-engravers, photographers, and lithographers.

For pasting paper and other paper products, dextrine pastes are very widely used. Starch pastes are used also for this purpose.

EXPERIMENT 377—How to Make Starch Paste.



Put about 10 measures of powdered starch in a test tube, then fill the tube ¼ full of water. Shake the tube thoroughly so that the starch forms a cream with the water.

Fill another test tube half full of water and heat it to boiling. Now pour the starch cream into the test tube, containing the hot water and then boil the mixture for just a few moments longer. When cool, you will have a thick starch paste which is fine for pasting on labels or for use with paper of any sort.

If you desire to keep the starch paste, add a drop or two of carbolic acid or other preservative, as mold will soon form if you do not take this precaution.

EXPERIMENT 378—How to Make Flour Paste.

Paste made from wheat flour is used in various kinds of light work. This paste can be made in the following way: Place in a test tube 10 measures of white flour, fill the tube half full of water and then boil the solution over a flame for a few minutes. As soon as the flour forms a paste with the water, it is time to stop boiling, as too much boiling injures the paste. If it is desired to keep the paste for any length of time, it will be necessary to add a little preservative, as otherwise it will mold.

For mending glass and chinaware, sodium silicate solution (water glass) provides an excellent cement. It is widely used also in making corrugated packing cases.

EXPERIMENT 379—Sodium Silicate Cement.

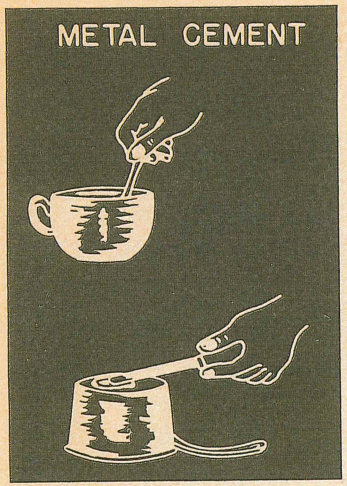
To mend a piece of broken glass, paint both the broken surfaces with a thin coat of Sodium Silicate Solution (No. 19), using a soft brush. Allow the solution to dry for a few moments and then press the edges firmly together. Wipe off any of the adhesive which squeezes out of the joint with a damp cloth, and then allow the mend to harden for at least 24 hours. It will make a strong, nearly transparent mend which will hold well, although it is not waterproof. Sodium silicate solution is one of the most popular cements for mending glass.

EXPERIMENT 380—Cement for Blue China.

In case you have a piece of blue china or porcelain to mend, mix 2 measures of Azurite (No. 39) with a few drops of Sodium Silicate Solution (No. 19) and apply this paste to the surface to be joined. This forms a blue cement which can be made to match the article to be mended.

EXPERIMENT 381—Making a Cement from Lime and Egg White.

A cement which is very easy to make and which is highly recommended for mending crockery ware is made with lime and white of an egg. To make this cement, put a small quantity of the white of an uncooked egg in a cup and add enough Calcium Oxide (No. 20) to make a good paste. Stir well and apply this paste quickly to the edges and then press firmly together and the cement will very soon set and become strong. Mix only enough of this cement for your immediate needs because it hardens very quickly and can only be used when freshly mixed.



EXPERIMENT 382 — A Cement for Mending Metal Vessels.

Place in a cup 5 measures of powdered iron, 1 measure of Ammonium Chloride (No. 9), and 1 measure of Sulfur (No. 1). Add 2 or 3 drops of water to this mixture and stir it until it has formed a thick paste. Now fill the hole or crack which you wish to mend with this mixture, forcing it in tightly. Allow the mass to harden for about 24 hours and then heat the pot or pan over the stove for a little while. This sets the cement so that it forms a good strong mend.

EXPERIMENT 383 — How to Make "Rubber Cement".

If you have some carbon tetrachloride or benzine cleaning fluid, you can easily make "rubber cement." Cut a few scraps from a crepe rubber shoe sole and place them in a dry test tube. Add ½ test tube of carbon tetrachloride or benzine and then close the mouth of the test tube with your thumb. Shake the test tube vigorously for a minute, place a stopper in it and set it aside. After an hour or so, notice how the scraps of rubber begin to swell. Shake the test tube occasionally, and soon the pieces will be swollen to several times their original size and their outline will be very faint. Keep the test tube closed with a cork or a plug of cotton, so that the solvent will not evaporate. After several days the rubber will be completely dispersed through the liquid and you will have a fairly thick, sticky solution. Save this rubber cement for future use. Adding a small amount of powdered rosin will increase the adhesive power of the cement.

Vulcanized rubber does not dissolve readily, but crepe shoe soles are usually unvulcanized.

SOAPS AND DETERGENTS

The manufacture of soap goes back to ancient times. Its discovery has been credited to the Gauls and a complete factory for the manufacture of soap has been unearthed in the ruins of Pompeii. Soap was probably first manufactured in England around 1650 and for many years it was considered a luxury. Today, of course, the use of soap in our country is universal. It is estimated that each person in the United States uses several pounds of soap each year. Soap is used not only in the household and in laundries, but large quantities are also used by various industries.

The manufacture of soap is the result of a chemical reaction. Fats, such as lard, cottonseed oil, and linseed oil are compounds of glycerine with what are known as "fatty acids." If fat and oils are treated with an alkali, the glycerine is set free and a compound of the alkali or base and the fatty acid is formed, which is soap.

There are hundreds of different types of soap, ranging from highly perfumed soaps manufactured from costly oils to common yellow laundry soaps made from waste fats.

EXPERIMENT 384—How to Make Soap.

Place 2 measures of Sodium Carbonate (No. 4) and 3 measures of Calcium Oxide (No. 20) in a test tube. Fill ⅓ full of water and boil for two or three minutes. Keep your face away from the mouth of the tube in case the liquid should spurt out while boiling.

Allow the solids to settle and pour off the clear liquid into another test tube.

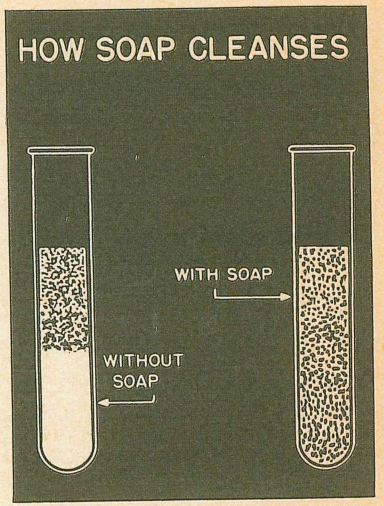
Add a lump of lard or butter about the size of a small pea and heat the liquid again for a few minutes. The fat will melt, then dissolve in the liquid, forming soap. Pour out a few drops on your hand and rub it; notice the soapy feel.

The soap can be precipitated from the liquid by adding 6 measures of common table salt and shaking it for a few minutes. Let stand and the soap will form a layer on top. The liquid underneath contains dissolved in it the glycerine which was in the fat.

EXPERIMENT 385—How Soap Cleanses.

Have you ever wondered why it is that soap will immediately remove dirt from your hands, while if you try to wash them without soap, it takes so much harder rubbing? Soap cleanses by reducing the surface tension of the water so that the dirt is thoroughly wetted. At the same time the lather formed by the soap holds the dirt in suspension, so that by rinsing away the soap lather, the dirt goes with it.

Obtain a small amount of kerosene or light lubricating oil and fill a test tube about ¼ full with this substance. Now add water until the tube is half full and shake vigorously for a few minutes until the oil and water are thoroughly mixed. When you stop shaking the oil and water separate very rapidly. Now add a small amount of soap solution to the oil and water in the test tube and shake again. This time the oil and water stay mixed together for some time before they separate and it is an easy matter to remove the oil while it is held by the soap and water.



EXPERIMENT 386—Testing Soap for Free Alkali.

With a clean knife, cut a fresh surface on a piece of dry soap and allow a single drop of Phenolphthalein Solution (No. 22) to fall on the exposed surface. If a red color appears where the drop of phenolphthalein is in contact with the soap, it shows that free alkali is present. Try this experiment with laundry soap and also with toilet soap. Laundry soap usually contains free alkali. Toilet soap should contain little or no free alkali, as this causes chapping and roughness of the skin. This is a very delicate test.

EXPERIMENT 387—Hydrolysis of Soap.

Put a small piece of soap in a dry test tube and add 2 or 3 drops of Phenolphthalein Solution (No. 22). Notice that there is only a faint pink color formed. Now fill the tube about half full of water and shake it thoroughly. You will find the solution becomes red, showing that an alkali was formed when the water was added to the soap. Water hydrolyzes soap and liberates a certain amount of free alkali even though the soap was neutral before the water was added.

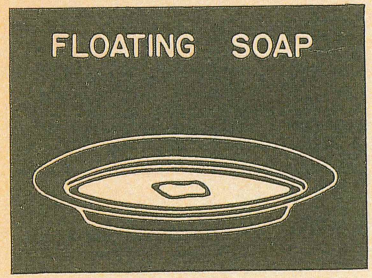
Household soaps usually contain added substances to increase their cleansing power, to improve their appearance or odor, or to give them various special properties. The following experiments illustrate a few of the commonly added materials.

EXPERIMENT 388—Transparent Soap.

Take a piece of white soap about the size of a small marble and put it in the spoon together with 4 or 5 drops of glycerine. Heat the spoon over the flame for a few minutes until the soap has melted and mixed with the glycerine. Upon cooling you will find that the mass is transparent. Transparent soaps are also made by mixing alcohol or sugar with ordinary soaps.

EXPERIMENT 389—How to Make a Colored Soap.

Put a heaping spoonful of soap shavings in a cup and add a few drops of colored ink and 3 or 4 drops of water. Grind the mixture thoroughly, and when the whole mass is a uniform color, work it up into a cake with your fingers. A good way of forming it into a cake is to press it into the bowl of the spoon, then remove it from the spoon and let it dry for a few hours. By using different dyestuffs almost any desired color of soap can be made.



EXPERIMENT 390—Making a Soap that Floats.

Put about 2 heaping spoonfuls of soap shavings in a dish and grind them up with enough water to form a stiff paste. Stir this paste well with your spoon and then scrape the whole mass into a pile on a plate or dish and allow it to dry over night. If you whip the soap and water thoroughly, you will find the next day that your mass will float even though the soap with which you started was not of the floating variety. The secret of making floating soap is to mix the soap, when it is soft, with air so that tiny bubbles of air are enclosed all through the soap, making it light and porous.

EXPERIMENT 391—Perfumed Soap.

Put a heaping spoonful of soap shavings into a dish and add 4 or 5 drops of perfume. Grind the mixture thoroughly and then form it into a cake by pressing it into the bowl of your spoon. Of course, various perfumes can be used to make any odor desired.

EXPERIMENT 392—How to Make a Medicated Soap.

These soaps are mixed with various antiseptic substances, such as carbolic acid. When used, a small amount of the antiseptic substance dissolves and comes in contact with the surface to which the soap is applied.

You can make a small quantity of medicated soap by mixing a little soap and boric acid with some water and grinding the mass in a dish until it is well mixed, and then drying the paste.

EXPERIMENT 393—Making Laundry Soap.

Put 3 or 4 measures of soap shavings in a dish and grind them to a paste with 4 or 5 drops of Sodium Silicate Solution (No. 19). When you have a good mixture scrape it in a pile on a plate or tin pan and let it dry.

Laundry soap frequently contains sodium silicate, as well as sodium carbonate. These materials aid somewhat in heavy washing, such as that carried on in commercial laundries.

EXPERIMENT 394—How to Make Borax Soap.

Soap is sometimes mixed with borax and used as a laundry soap. Grind together in a dish 5 or 6 measures of soap shavings with 2 measures of sodium borate (borax) and 3 or 4 drops of water. Scrape the paste onto a tin pan after it has been thoroughly mixed and then let it dry.

EXPERIMENT 395—How to Make Washing Powder.

Obtain a cake of white soap and shave a quantity of it into fine shavings with a knife. Put 2 measures of this soap in a dish together with 4 measures of Sodium Carbonate (No. 4) and 3 or 4 drops of water. Grind this mixture thoroughly and then scrape it into a bowl or onto a tin plate where it can dry.

The sodium carbonate will take up water of hydration, changing into washing soda, and when dry, the whole mass can be pulverized, and will be a sample of regular washing powder.

EXPERIMENT 396—Preparation of a Liquid Soap.

Put a spoonful of soap shavings into a dish and add water, a few drops at a time, stirring the mixture continually. When the proper consistency has been reached, pour the liquid soap into a bottle if you wish to keep it.

Since soap is usually the sodium salt of a fatty acid, we would expect that soluble salts of the heavier metals would produce a precipitate of an insoluble soap. That this is true is shown in the following experiments. The insoluble soaps of various metals find wide application in industry in making lubricating greases, in waterproofing cloth, in cosmetics, in paint and varnish as dryers and as dry cleaning soaps.

EXPERIMENT 397—Cobalt Soap.

Put 1 measure of soap shavings in a test tube half full of water and shake for a few minutes until the soap is entirely dissolved. Now add 2 or 3 drops of Cobalt Chloride Solution (No. 27). You will observe a very pretty pink precipitate, which is cobalt soap.

EXPERIMENT 398—Iron Soap.

Prepare a solution of Ferric Ammonium Sulfate (No. 21) by dissolving 1 measure of this substance in a test tube half full of water. Add a few drops of soap solution and you will obtain a thick brown precipitate of iron soap. Allow this precipitate to settle, pour off the clear liquid and empty the solid matter on a piece of white paper. Notice the rustlike stain which this substance makes. This is what happens when clothes are washed with soap in water containing dissolved iron compounds.

EXPERIMENT 399—Aluminum Soap.

Dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube half full of water. Put 1 measure of soap shavings in another test tube, fill half full of water and shake for a few minutes until the soap is dissolved. Now add a few drops of the soap solution to the aluminum sulfate solution and a thick white precipitate of aluminum soap will be formed.

EXPERIMENT 400—Strontium Soap.

Dissolve 1 measure of Strontium Chloride (No. 10) in a test tube half full of water. Add a few drops of soap solution and examine the thick white precipitate of strontium soap which forms.

EXPERIMENT 401—Zinc Soap.

Dissolve a small piece of zinc metal in a test tube ¼ full of water to which has been added 2 measures of Sodium Bisulfate (No. 7). Warming the solution will make the zinc dissolve more rapidly. Pour some of the clear solution into a clean test tube and after the liquid has become cool, add a few drops of soap solution. A thick precipitate of white zinc soap will result.

EXPERIMENT 402—Copper Soap.

Put 1 measure of Azurite (No. 39) and 1 measure of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Shake the test tube until the solids are entirely dissolved. Now add a few drops of soap solution and a blue precipitate of copper soap will result.

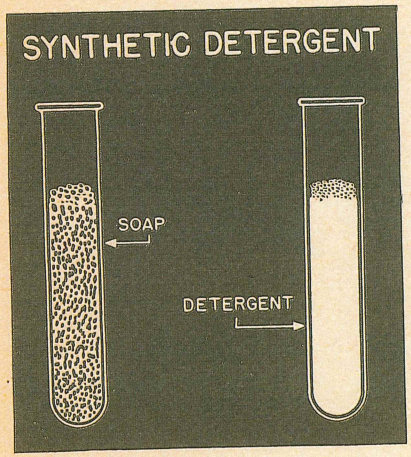
EXPERIMENT 403—Magnesium Soap.

Dissolve 1 measure of Epsom salts (magnesium sulfate) in a test tube half full of water. Add a few drops of soap solution and a white precipitate of magnesium soap will form.

SOAPLESS SOAPS

A new class of cleansing agents known generally as synthetic detergents has recently come into widespread use. These substances, while used for the same purposes as soap, are chemically of quite a different nature. Many synthetic detergents are based on a class of compounds known as sulphonated oils, which, like soaps, have the cleansing properties due in part to their reducing surface tension of water. Synthetic detergents have certain advantages over soap for certain uses, because when used with hard water, they do not leave a film or residue of insoluble calcium or magnesium soap.

EXPERIMENT 404—Synthetic Detergents.



If your water supply is somewhat hard, that is, if it contains dissolved calcium or magnesium salts, this will prove to be an interesting experiment. Fill a test tube half full of water, add half a measure of a synthetic detergent, such as Dreft. If your synthetic detergent is in liquid form, add just a drop or two. Shake the tube until the solid is dissolved and then set it aside for a few minutes. In another test tube half full of hard water, place half a measure of soap shavings. Shake this test tube until the soap is all dissolved and then set the second tube beside the first one. If the water which you used is fairly hard, you will notice a slight cloudiness in the test tube in which you dissolved the soap, whereas the test tube containing the synthetic detergent will be perfectly clear. This cloudiness in the soap solution is due to a precipitate of calcium or magnesium soap.

THE CHEMISTRY OF LEATHER TANNING

Most leather is made from the skins of steers and calves. Small quantities of special leather are made from the skins of other animals, including pigs, sheep and horses. The tanner obtains his skins or hides from meat packing houses, where the animals are slaughtered. The hides are first treated with sodium chloride, which acts as a dehydrating agent and as a preservative, thus keeping them from molding until they are to be tanned.

There are two general processes of tanning — bark tanning and chrome tanning. In bark tanning, oak bark is probably most widely used, although hemlock bark, chestnut bark and in some cases, extracted tannic acid or tannin may be used.

The first step in the tanning of hides by the bark method is to remove the hair from the hides. This is done by soaking the hides in tanks containing lime and sodium sulfide solution, which loosens the hair which is completely removed by scraping. If concentrated de-hairing solutions are used, the hair may be completely dissolved.

The next step is called bating, in which the hides are treated with certain enzymes which make them soft and flexible and loosen their texture so as to allow better penetration of the tanning liquor.

After bating, the hides go to the tanning yard, which consists of a series of tanks containing tanning liquors. These liquors are made by digesting the bark or other source of tannic acid with hot water and steam until the required concentration is obtained. The hides are first placed in a weak tanning liquor and each day are transferred to a liquor of higher concentration.

After about 15 days of soaking in the tanning vats, the hides are thoroughly impregnated with tannic acid, which combines with the protein of the hide fibers, increasing their toughness and resistance to water and preventing putrefaction. After being washed and passed through giant ringers, the hides are split by machines into sheets of leather. Thicknesses down to 1/64 of an inch may be produced. The leather from the outside of the hide, next to the hair, is considered to be the best and is called "top grain" leather.

The leather now goes to the dye vats where it can be given any desired color. It is then dried and stored in rooms having carefully controlled humidity so that the leather will absorb the proper amount of moisture. After stretching and oiling, the leather goes to the finishing department, where it may be treated with shellac, gelatin, casein, albumen, wax, lacquers or varnishes, depending upon the products to be manufactured from it.

The tanning of leather by the chrome process requires much less time than bark tanning and the resultant product has a greater resistance to moisture and heat. In chrome tanning, the hides are cleaned and are immediately placed in a pickling or tanning solution which contains salts of the metal chromium. Every tannery has its own special formula, but in general, the chrome tanning solution consists of sodium or potassium dichromate, sulfuric acid and glucose (corn syrup). The hides are allowed to remain in the chrome tanning solution for 24 hours, being continually agitated by huge paddles. In the chrome tanning process, the hair is dissolved from the hides by the tanning solution. When the hides are removed from this solution, they are treated with soda to neutralize the acidity. Chrome tanning may be done in as short a time as three or four hours under very favorable conditions.

The hides are now split, dyed and finished in very much the same manner as that for bark tanned leather. To produce white chrome leather, the hides are treated with titanium dioxide, a white pigment which whitens the fibers of the leather. To produce imitations of such leather as pigskin, walrus hide and reptile skin, etc., the split leather is placed in giant presses where it is embossed with the required design at temperatures of about 180° F. and under pressures of many tons. Patent leather is made by spraying the leather with synthetic lacquers.

EXPERIMENT 405—The Chemistry of Tanning.

Dissolve a few drops of the white of an egg in half a test tube of cold water. In another test tube, dissolve 1 measure of Tannic Acid (No. 15) in half a test tube of water. Add the solution of tannic acid, a few drops at a time, to the solution of the white of egg and notice the precipitate which is formed. Skins of animals contain a large amount of albuminous substance similar to egg-white, and when they are exposed to a solution containing tannic acid, a tough, pliable mass (leather) results. If hides and skins were not tanned they could not be employed for the purposes for which leather is used.

Many plants and their fruits contain considerable amounts of tannic acid, which may easily be detected by the intense black color produced by the reaction of tannic acid with a ferric salt.

EXPERIMENT 406—Testing Tea for Tannic Acid.

You can either boil a few tea leaves in a test tube or wait until your mother makes some tea and then test a little of that with ¼ measure of Ferric Ammonium Sulfate (No. 21). You will find this results in a beautiful black color, showing that there is tannic acid in tea.

EXPERIMENT 407—Testing Persimmons for Tannic Acid.

The green persimmon contains quite a little tannic acid, which perhaps accounts for the way in which it puckers up our mouths when we try to eat it. If you live near a persimmon tree, get a nice green one, cut a section and boil it for a few minutes in a test tube ½ full of water. Test this solution with ¼ measure of Ferric Ammonium Sulfate (No. 21), and the black color formed will show that there is tannic acid present.

EXPERIMENT 408—Testing Hemlock Bark for Tannic Acid.

If you can obtain small pieces of hemlock bark, cover about half a test tube full of very fine shavings with water and heat for a few minutes. Now pour the solution in another test tube, being careful to leave the pieces of bark in the first test tube.

To the solution in the second test tube, add about ¼ measure of Ferric Ammonium Sulfate (No. 21). The intense black color which is formed will show that you have tannic acid.

EXPERIMENT 409—Testing Oak Bark for Tannic Acid.

Cut some fine shavings from a piece of oak bark. Do not use the outside of the bark, but cut the shavings from the inside. Fill a test tube half full of these shavings, add water and heat for 4 or 5 minutes. Pour the solution into another test tube and test with ¼ measure of Ferric Ammonium Sulfate (No. 21). You will again obtain a black color, showing that oak bark contains tannic acid.

EXPERIMENT 410—Testing Chestnut Bark for Tannic Acid.

An extract of tannin which is used by tanners is often obtained from chestnut bark. If you can obtain a small piece of chestnut bark, fill a test tube half full of shavings, add water and boil for some little time. Pour the solution into another test tube and test it with ¼ measure of Ferric Ammonium Sulfate (No. 21). A black color will show the presence of tannic acid.

EXPERIMENT 411—Testing Sumac Wood for Tannic Acid.

Sumac wood contains quite a little tannic acid. Boil a few chips for a little while and then test the solution as in the preceding experiments.

Tannic acid extract was sometimes made from sumac wood, but this source is seldom used at present.

Salt of some of the metals gives precipitates with tannic acid. The precipitates are called tannates, although they are not definite chemical compounds.

EXPERIMENT 412—Formation of Ferric Tannate.

Dissolve ½ measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. In another test tube about half full of water, dissolve ½ measure of Tannic Acid (No. 15). Pour the contents of the first test tube into the second and note that a black precipitate forms at once.

EXPERIMENT 413—Formation of Cobalt Tannate.

Put 2 drops of Cobalt Chloride Solution (No. 27) in a test tube ¼ full of water. Add a few drops of tannic acid solution and notice that a brown precipitate gradually forms, becoming heavier as the liquid stands.

EXPERIMENT 414—Formation of Copper Tannate.

Put ½ measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7) in a test tube half full of water and shake the tube to thoroughly dissolve the solids. In another test tube half full of water, place 1 measure of Tannic Acid (No. 15). Pour this solution into the first test tube and notice the bluish brown precipitate which forms. This precipitate will gradually become heavier.

RUBBER

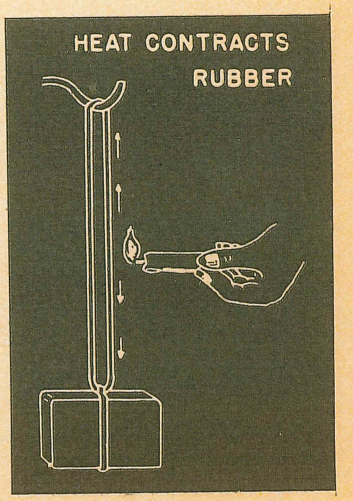
Rubber is one of the most versatile substances known. It can be made as hard as a rock or as soft as a sponge cake; as delicate as tissue paper, or tough enough to roll 40,000 miles underneath a heavy load. This interesting material has become an important part of our daily lives and it is essential to our modern civilization.

Natural rubber comes from the sap of a tropical tree, while synthetic rubber is made from organic chemicals derived from petroleum and other sources. Rubber is known chemically as a hydrocarbon, that is, a compound of carbon and hydrogen. The carbon atoms in a molecule of rubber are joined in long chains which lie in kinks and coils. When the rubber is stretched, these chains of atoms are more or less straightened out, but they tend to curl up again as soon as released. This peculiar molecular structure of rubber perhaps accounts for the fact that, unlike almost all other substances, stretched rubber contracts when it is heated.

EXPERIMENT 415—Contraction of Rubber When Heated.

Hang a stout rubber band on a hook so that it hangs clear of the wall and from it suspend a weight heavy enough to stretch it to about three times its original length.

Now warm the band by moving a candle flame very cautiously up and down alongside the rubber. You will have to be very careful not to heat it so hot that it softens and breaks. A better way to heat it is with a radiant electric heater. Watch the weight closely as you warm the rubber and you will see it slowly rise through a fraction of an inch. As the rubber cools after you remove the heat, the weight will again descend. The contraction of the rubber is, of course, comparatively slight, but it can be seen disinctly if the experiment is carefully done.



EXPERIMENT 416—Showing that Rubber Gets Warm when Stretched.

Take hold of a good wide rubber band with your hands about half an inch apart and stretch it vigorously a few times. Now with the rubber relaxed, bring it against your lip, where the skin is sensitive to heat, and stretch it quickly nearly to its limit. Do you feel the rubber get warm? Still holding it gently against your lip, let it contract again and you will feel it become cool.

While we are on the subject of rubber, we ought to mention one of its humble cousins, chicle. Chicle is obtained by boiling the milky sap of a tree which grows in Mexico and Honduras. Several thousand tons of chicle are imported every year for use in chewing gum, although synthetic substitutes are now being used more and more.

EXPERIMENT 417—Obtaining Chicle from Chewing Gum.

Most chewing gum contains about 40% chicle, the remainder being chiefly sugar and flavoring, though there is a tendency to substitute synthetic plastic materials more and more for chicle.

Chew a stick of your favorite gum for 15 or 20 minutes to extract all of the soluble materials and then rinse it in cold water and save it for the following experiment. If you simply soak a piece of unchewed gum in cold water, the sugar will dissolve out, of course, in time, but the chicle will not gather together into a single mass as it does when you chew the gum. Instead it remains as little particles scattered through the water.

EXPERIMENT 418—Effect of Heat and Cold On Chicle.

Chicle, although very similar to rubber in its chemical composition, is much more easily affected by heat than raw rubber. Soak your piece of chicle obtained in the preceding experiment in ice water for a few minutes, and when it is thoroughly chilled, notice how it has lost its stickiness and become hard and fairly stiff.

Now remove it from the ice water and let it warm up, putting it on a piece of paper near a radiator or stove. You will not have to warm it very much before it becomes extremely sticky on the surface and loses its stiffness. The synthetic chewing gum bases are affected by heat and cold in very much the same way as is chicle.

Let the chicle stand exposed to the air for a day or two, and you will find that it turns a little darker and becomes harder and less sticky on the surface. This is due to a chemical reaction with the oxygen of the air. It is very similar to the change which occurs in rubber when it ages and becomes weak and brittle, though fortunately with rubber, the change is much slower.

GEMS

The diamond is the prince of gems. It is a crystallized form of carbon and differs from black, flaky graphite only in the arrangement of atoms in the crystal. The greater portion of the world's production of uncut diamonds comes from South Africa. Since the diamond is the hardest known substance, cracked and defective specimens are used in grinding and polishing operations and in tools for drilling.

Very tiny diamonds of no commercial value have been made artifically. Melted iron is heated to a very high temperature in an electric furnace and pure charcoal is dissolved in it. The white hot liquid iron is then suddenly cooled. It solidifies first on the surface, causing a very high pressure in the interior. Under these conditions some of the carbon crystallizes into microscopic diamonds. The imitation diamonds commonly sold are made of glass.

Rubies are composed of aluminum oxide colored by various impurities. They are made artificially on a commercial scale by allowing powdered aluminum oxide to fall through an oxy-hydrogen flame onto a little pedestal. The powder melts in passing through the flame and builds up a pear-shaped mount of fused material. Traces of chromium compounds are added to give the red color. Sapphires are made in the same way but are colored blue with cobalt oxide.

Emeralds are composed of beryllium aluminum silicate, colored by minute amounts of chromium. Perfect emeralds are among the most costly gems.

EXPERIMENT 419—Making Imitation Rubies.

With your glass rod, place a drop of Sodium Silicate Solution (No. 19) on a sheet of glazed paper. Let it stand for about 15 minutes to stiffen a little and then touch it several times with a pin-head wet with Phenolphthalein Solution (No. 22). Let it harden overnight, and you will have a chemical ruby. To make it thicker, you can add another drop of sodium silicate solution.

Pearls are composed of about 75% calcium carbonate, the same thing chemically as limestone and marble, 20% is water, and the remainder is a horny protein material. They are formed when an irritating particle lodges inside the shell of an oyster. Often tiny grains of sand are intentionally implanted in oyster shells to induce the formation of pearls. You have heard the story of Cleopatra, who made a rare and expensive drink by dissolving pearls in vinegar. Pearls, being largely calcium carbonate, are soluble in dilute acids, such as vinegar, and carbon dioxide gas is given off, producing an effervescence. Cleopatra's drink was a solution of calcium acetate. It could not have been very pleasant to drink.

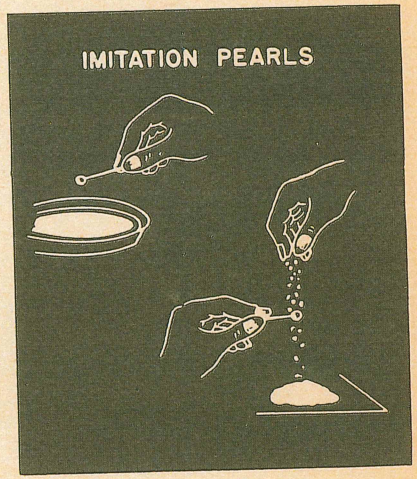
Imitation pearls are made in various ways, but man has never succeeded in making artificial pearls that equal those of the oyster.

EXPERIMENT 420 — Imitation Pearls.

Prepare a mixture in your spoon of 2 measures of powdered chalk in 2 or 3 drops of water and add 10 drops of Sodium Silicate Solution (No. 19). Mix thoroughly.

Obtain a small wooden bead without a hole and stick it on the end of a pin. The pearl is to be built up by repeated dippings into the mixture. Each time you dip the bead into the mixture, let it dry for 5 minutes before dipping again. After you have dipped it about five times, dust the pearl with a little powdered chalk and see that it is spread evenly over the entire surface. When it dries, continue dipping it into the solution, and about each fifth dip, sprinkle a little powdered chalk on it.

When the pearl is nearly the desired size, add a drop of Phenolphthalein Solution (No. 22) to the mixture in the spoon. This will give a pink tinge to your pearl.



PHOTOGRAPHY

Photography is the art by which the image of an object is formed by lenses on a sensitive film and made permanent by certain chemical processes. The effect of light in changing the nature of certain chemical substances was known to the ancients, but the first practical use of these reactions was made in 1802 when Thomas Wedgwood produced sun pictures on white paper treated with a silver salt.

In modern photography, the sensitive plate or film is usually coated with silver bromide and silver chloride held in a layer of gelatin. When you

click the shutter of your camera, light from the object which you are photographing passes through the lens of the camera and strikes the sensitized plate or film. This light changes the silver bromide and chloride in such a way that the developer can act upon them. The film or plate is now put into a developing bath which consists of a solution of pyro-gallol or a similar organic compound. The silver bromide and chlorine grains which have been acted upon by light are reduced to metallic silver by the action of the developer. The next step is to immerse the plate or film in a bath of "hypo," or sodium thiosulfate, which dissolves out all of the silver bromide which was not affected by the light. Hence, the plate or film now has dark spots of metallic silver where the light originally struck. This plate or film is a reverse picture or "negative."

To obtain a photographic print the negative is placed over a sheet of sensitized paper and light allowed to shine through it. This time the silver bromide is reduced under those parts of the negative which are light. The paper is then developed and fixed in practically the same manner as the negative was, and you have your finished photographic "print."

A film or negative can be printed on blueprint paper instead of on the regular silver papers. Papers treated with a solution containing ferric ammonium citrate and potassium ferricyanide become blue when exposed to light, owing to the reduction of the ferricyanide and the formation of Prussian blue. Blueprint paper is used more extensively for printing mechanical drawings, plans, and maps than it is for printing pictures from film negatives.

EXPERIMENT 421—How to Make a Blueprint.

Obtain a small piece of blueprint paper from a photographic supply store. Be careful not to expose it to strong light and when you open the package to cut off pieces, work in a semi-dark room or under a red light. Lay a small key, or a piece of transparent glassine paper on which you have written with black ink, on the surface of the piece of blueprint paper. Use a paper clip to hold the object firmly against the surface of the blueprint paper and expose the paper and object to bright sunlight for 2 minutes. Remove the object from the paper and quickly immerse the blueprint paper in water to develop your print. Now dry the blueprint and you have a print of the outline of the object which you held over it. If you have some pieces of developed film or developed photographic plates, these can be printed on the blueprint paper, and with a little experience, you can obtain excellent prints. Drawings or pictures made on semi-transparent paper in black ink can be reproduced in this manner on blueprint paper.

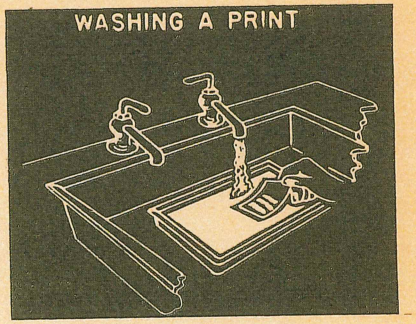
EXPERIMENT 422—Making a Photographic Print.

Obtain from a photographic supply store a few sheets of "proof paper." In a dimly lighted room, place a negative (or a piece of glassine paper on which you have drawn a design with black ink) against the sensitized side of the proof paper, cut slightly larger than the negative. Place the paper and negative on a flat surface and cover them with a piece of clean glass. Expose the negative to strong daylight for from 1 to 3 minutes, depending upon the density of your negative.

When the paper beyond the edge of the negative has turned a dark brown color, remove the glass and negative and examine your photographic print. The silver halides—silver bromide, silver iodide and silver chloride—have been darkened by the light which reached them through the negative, and of course they have been darkened most strongly in the places where the most light reached them, that is, underneath the light parts of the negative. The photographic print which you have just made will gradually darken unless something is done to prevent the light from acting further upon it. This is done by the process of "fixing."

EXPERIMENT 423—How to "Fix" a Photographic Print with "Hypo."

Place in a glass 4 teaspoonfuls (about 1 ounce) of water, and add 15 measures of Sodium Thiosulfate (No. 16). Stir this solution with a glass rod until the chemical is completely dissolved. Immerse your photographic print in the "hypo" solution just made until it is completely covered. Allow it to remain in the "fixing bath" for about 5 minutes. By the end of this period the unchanged silver compounds in the paper will be dissolved out, and light will no longer have any effect upon the print.



After fixing your print by immersing it in the fixing bath, it must be carefully washed in running water for approximately 10 minutes to dissolve all of the "hypo" which remained in the fibers of the paper.

This can best be done in a shallow dish or pan placed in a sink under a stream of water from a faucet. Occasionally turn the print so that it is washed on both sides. Do not hasten this operation, as a poorly washed print will eventually show yellow stains which are highly objectionable.

Upon completing the washing operation, hang the print up by one corner, using a paper clip, so that air can get to it freely from all sides.

The dark material which forms the image in an ordinary photograph is silver in the form of tiny particles. By treating the print with various reagents, the silver can be converted into compounds or replaced by other metals. This changes to some extent the color of the print. The operation is called toning.

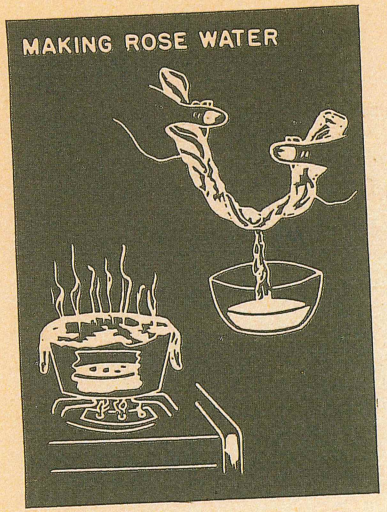
EXPERIMENT 424—How to Tone a Photograph.

In a test tube half full of water, place 1 measure of Sulfur (No. 1) and 1 measure of Calcium Oxide (No. 20). Shake thoroughly and then boil for a couple of minutes. Pour the liquid into a saucer and immerse in it an ordinary black-and-white photographic print. Let it soak for about half an hour and then remove it and wash it carefully and thoroughly in running water. After it has dried, you will have a sepia-toned print. The silver has been converted into silver sulfide.

PERFUMES AND ESSENTIAL OILS

Since the beginning of history the sweet odors of flowers, certain woods, leaves, and roots have been highly prized by man, and the art of perfumery is of ancient origin. Each year millions upon millions of dollars are spent for the products of the perfume and essential oils industries.

An essential oil is one which contains the "essence" of the flower fruit odor. Years ago all odor concentrates and essential oils were obtained directly from their natural sources, but today many of them are produced synthetically by the chemist. Many flavoring extracts, perfumes, and the perfuming agents used in cosmetics are products of the chemical laboratory. These materials with pleasant odors are organic compounds and although of a highly complex nature, the chemist has been able to produce synthetic products more cheaply than the natural products can be extracted.



EXPERIMENT 425—How to Make Rose Water.

Obtain a small quantity of rose petals during the rose season and put them in a sauce pan. Add enough water to cover the petals.

Gradually bring the water to boiling. When it starts to boil, cover the pan with a piece of clean absorbent cloth and as soon as this cloth has become wet from the steam wring it into a bowl and then put the cloth over the pan again and catch more steam. The fragrance of the rose petals distils over with the steam so that by condensing the steam in the cloth and collecting the water we can obtain the sweet oil of roses. After you have collected rose water by this method, smell and notice the sweet rose odor which it has.

EXPERIMENT 426—How to Make a Sachet Jar.

Mix together in a cup the following spices or any of them which you can obtain: Half a teaspoonful of cloves, half a teaspoonful of allspice, half a teaspoonful of cinnamon and about half a teaspoonful of extract of vanilla. Stir these materials together until they are thoroughly mixed.

Now obtain about a quart of dried rose petals and mix this spice which you have just prepared with the rose petals. When it is thoroughly mixed, you will find that it gives a very sweet, pleasant odor, and it can be used in sachets or put in jars, thus making rose jars. Whenever the cover is removed from these jars, the sweet odor will fill the room.

EXPERIMENT 427—How to Make a Sage Sachet.

One of the common plants which has a very sweet odor is the sage. The leaves of this plant when dried retain their odor for a long time, and they can be used to good advantage in sachets and other places where a sweet odor is desired.

EXPERIMENT 428—Extracting Orange Oil.

Remove the outside yellow peel of an orange and crush this in a bowl. The fragrant oil which you obtain is oil of orange.

EXPERIMENT 429—Extracting Oil of Lemon.

Remove the thin outside yellow peel from a lemon by means of a sharp knife. Place this peel in a bowl and crush it until you have pressed out the oil. Pour this oil into a test tube or bottle and notice its fragrance and taste. This is oil of lemon.

EXPERIMENT 430—Preparation of Oil of Sassafrass.

If you can obtain some dried sassafrass roots, strip the bark off and place the clean roots in a small bottle. Add about 2 teaspoonfuls of rubbing alcohol and allow the bottle to stand for a day or two, shaking it every hour. At the end of this time, pour out several drops of your liquid onto a clean handkerchief and notice the odor of sassafrass. The solution in the bottle is an impure oil of sassafrass, extracted by the alcohol. Most essential oils are much more soluble in alcohol than in water and this is one of the methods used in obtaining costly natural oils.

EXPERIMENT 431—Oil of Cloves.

Place 3 cloves in a test tube and add water until the test tube is ½ full. Hold the test tube with the test tube holder, heat it over the flame of your alcohol lamp until it begins to boil, and continue to boil for about 3 minutes. Remove the test tube from the flame and when it is cool, place a drop of the solution on your tongue and notice the biting taste characteristic of cloves. The solution which you have prepared in your test tube contains an impure oil of cloves. Clove oil consists largely of a compound called eugenol. It is commonly used as "toothache drops."

EXPERIMENT 432—The Extraction of Anethole.

Anethole is an organic compound which is present in such oils as anise, fennel and pine. Obtain some dried anise roots from a drugstore and break off several pieces. Drop these into ¼ test tube of water and boil for several minutes. Remove the test tube from the flame and smell the vapors which are being evolved. You will at once notice the ethereal odor of the anethole which was present in the anise roots. This was extracted from the roots by boiling. The chemical name of anethole is para-allylphenyl methyl ether. It is often used to flavor toothpaste.

EXPERIMENT 433—The Preparation of Cymene.

Cymene, sometimes called para-methylisopropyl benzene, occurs naturally in essential oils, such as oil of thyme, oil of caraway, and oil of eucalyptus. Cymene can be obtained from any of these oils, or by the action of a dehydrating agent on camphor.

Obtain a small piece of gum camphor and chip off enough to make 2 measures. Place this in a dry test tube and add 2 measures of Sodium Bisulfate (No. 7). Now heat the tube gently over the flame of the alcohol lamp until vapors are evolved. Allow the tube to cool and then smell cautiously at the mouth. You should be able to detect the sharp, pleasant odor of cymene.

EXPERIMENT 434—Essential Oils from Rosin.

Rosin is the residue remaining upon the distillation of turpentine. It is a dark amber-colored material which melts at about the temperature of boiling water. It is widely used in the manufacture of varnishes, ointments and soaps, and in the paper industry.

Rosin is composed largely of abietic acid. When heated, it is separated into rosin spirit, a mixture of volatile substances used in varnishes, and into rosin oil, from which lubricating greases can be made.

To make some rosin oil, place 2 measures of rosin in a dry test tube and heat gently over a flame until a thick sticky mass is formed in the test tube and fumes are given off. Do not let the material char. Notice that part of the vapors condense on the upper part of the test tube, showing that the constituents of rosin vaporize at different temperatures.

EXPERIMENT 435—Essential Oils in Spruce Gum.

Spruce gum is a resin containing many volatile essential oils, especially oil of spruce.

Obtain some exuded resin from a spruce tree and place 2 measures of this material in a test tube. Heat gently over a flame, being careful that the material in the test tube does not become charred. Allow the tube to cool and smell at the mouth. This time you will detect an entirely different odor from that in the preceding experiment. It is due to the formation of oil of spruce.

ELECTROCHEMISTRY

The field of electrochemistry is growing quite rapidly and the future of this branch of chemistry is becoming of greater importance in our daily lives. The energy supplied by batteries, the separation of metals and non-metals from their ores or salts by electrolysis, electroplating and the relatively new field of controlled electron emission or electronics, all belong to the field of electrochemistry.

When a current of electricity is passed through certain solutions, a chemical reaction occurs. Likewise, when certain chemical reactions occur, a current of electricity is produced. Let us first see how we can make an electric current by chemical action. Then we shall go on to study the chemical effects produced by the current.

WET CELL BATTERIES

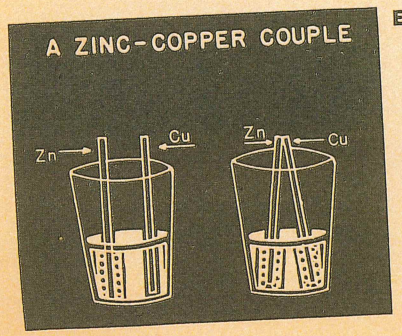
You have learned that very frequently when a chemical reaction takes place, heat is liberated. Heat is a form of energy, and electricity is also a form of energy, so it is not at all surprising that most reactions can be made to liberate electricity instead of heat. Theoretically, a current can be obtained from any chemical reaction in which there is oxidation and reduction, but in practice it is often impossible to get the current in sufficient amounts to be used. One of the simplest ways of generating a current is to dip a piece of zinc and a piece of copper into an acid.

EXPERIMENT 436—The Action of an Acid on Zinc.

Dissolve several measures of Sodium Bisulfate (No. 7) in a test tube full of water, and pour it into a glass or tumbler. Clean a zinc strip with fine sandpaper and place it in the liquid. Bubbles of hydrogen gas will be seen forming on the zinc. These bubbles indicate that a chemical reaction is taking place between the zinc and the sodium bisulfate solution which acts as an acid. The acid attacks the zinc and releases hydrogen.

EXPERIMENT 437—The Action of an Acid on Copper.

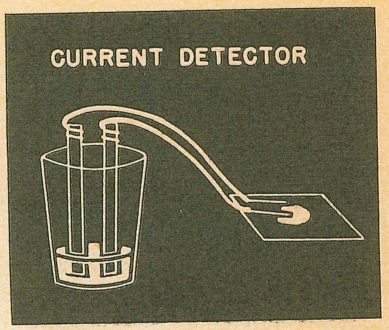
Remove the zinc strip in the above experiment and in place of it substitute a clean, bright copper strip. Notice that no bubbles are formed upon it, showing that the acid substance does not attack the copper.



EXPERIMENT 438—The Action of an Acid on a Zinc-Copper Couple.

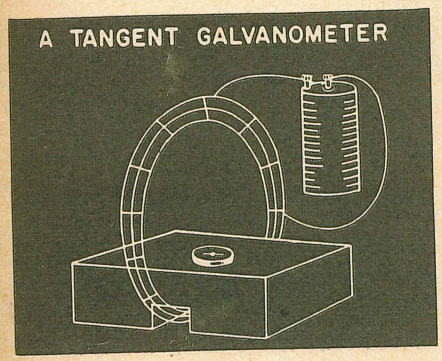
Now place both your copper and zinc strips in the sodium bisulfate solution without allowing them to touch each other. The bubbles still arise from the zinc, but not from the copper. Now touch the ends of the copper and zinc strips together outside of the liquid. You will see bubbles of gas being liberated at both the zinc and copper strips. Disconnect the metal and the bubbles of gas immediately cease to form on the copper strip. This shows that hydrogen is deposited upon the copper only when an electrical contact is made between the two strips.

EXPERIMENT 439—How to Produce an Electric Current.



Scrape both ends of two small copper wires until they are bright. Twist one end of one wire about a zinc strip and one end of the other copper wire around the end of a copper strip. Place the strips in the sodium bisulfate solution which you used in the last experiment. Do not allow the strips to touch each other. Place the free ends of the copper wires about ½ inch apart on piece of soft white paper 1 inch long and ½ inch wide. Wet the paper with several drops of salt water, made by dissolving several measures of table salt in ½ test tube of water. Also add to the paper between the wires, a drop of Phenolphthalein Solution (No. 22). In several seconds you will see that a pink color has formed on the wet paper, around the wire which is connected to the zinc strip. This is a visual test for the presence of electricity. The zinc and copper strips immersed in the sodium bisulfate solution are producing an electrical current. The wire about which the pink color formed is attached to the negative terminal of the cell, while the other wire is from the positive terminal. A pair of radio earphones will produce a loud click, indicating the presence of an electrical current, if you place the phone tips momentarily on the terminals of the cell you have made.

EXPERIMENT 440—Making a Tangent Galvanometer.



Since some of the experimental cells you will construct do not deliver a current sufficiently strong to give a test when using the phenolphthalein paper, you will need a sensitive testing device for further work.

Perhaps the most sensitive instrument which will detect an electric current and which is extremely easy to make is a tangent galvanometer. To make this instrument, wrap a coil of insulated wire about 4 or 5 inches in diameter and consisting of anywhere from 50 to 100 turns. Set the coil upright and hold it in this position with a wooden block that has a slot sawed in one side. Place a magnetic compass on the block and rotate the block and coil of wire until the coil sets in a vertical plane pointing north and south.

If even a very small current of electricity is sent through the coil, the needle will turn toward a position at right angles to the plane of the wire coil. Test your instrument by connecting the ends of the coil to the terminals or binding posts of a dry cell, even an old one will do. A swing of the needle indicates that a current is flowing through the instrument. Now that you are equipped with a tangent galvanometer, there are other types of wet cells which you can make and test.

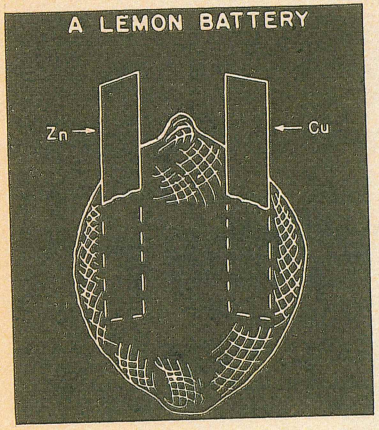
EXPERIMENT 441—Making a Sal-Ammoniac Cell.

Prepare a wet cell as before, only instead of using a sodium bisulfate solution, make the electrolyte or liquid of the cell by dissolving ½ teaspoonful of Ammonium Chloride (No. 9) in ½ glass of water. You will find that it too will produce a current. Such a cell is called a sal-ammoniac cell. Sal-ammoniac is another name for ammonium chloride. Test your cell using the tangent galvanometer.

EXPERIMENT 442—Other Types of Experimental Cells.

You can perform many experiments in the electrolytic production of electricity by using various other liquids as the electrolyte. The juices from fruits, vinegar, salt water, and other liquids can be tried. Your tangent galvanometer will be found to be more sensitive than the phenolphthalein paper. It will respond to weak electric currents which fail to cause the pink color when phenolphthalein-salt paper is used.

EXPERIMENT 443—A Lemon Battery.



Obtain a good juicy lemon and cut two small slits, one on each side. Now insert a clean strip of copper and a clean strip of zinc, as shown in the illustration. When you have your battery arranged, touch your tongue to the ends of the zinc and copper strips and you will be able to taste the slightly tingling sensation which shows the passage of an electric current.

Since zinc is a more active metal than copper, it has a greater tendency to dissolve in the citric acid of the lemon. The dissolving of the positively charged zinc ions leaves the zinc strip with a negative charge. Meanwhile the positive hydrogen ions of the citric acid are tending to deliver their positive charge to the copper strip.

Thus when a contact is made between the copper strip and the zinc strip, an electric current or flow of electricity occurs from the copper to the zinc.

Of course the lemon cell will not last very long nor will it generate very much current, but it serves to illustrate the general principles of the simple electric cell.

DRY CELL BATTERIES

A dry cell is not really dry. It is merely a wet cell which is non-spillable. If it were absolutely dry, it would produce no current.

You have seen in the experiment in which the zinc and copper strips were immersed in the sodium bisulfate solution that hydrogen bubbles formed around the copper strip. These bubbles prevent the current from flowing and therefore decrease the power of the battery. If it were possible to brush off the bubbles, the current would flow full strength as before. As brushing bubbles off every time they accumulate is impossible, they are usually eliminated by chemical means.

In the dry cell, this formation of bubbles, or polarization effect as it is called, is prevented by adding manganese dioxide to the electrolyte (liquid) of the cell. This compound is rich in oxygen. In the dry cell, ammonium chloride and zinc chloride in water solution are both used as electrolytes. Manganese dioxide oxidizes the hydrogen bubbles, forming water. The case of the cell is a cup of pure zinc, which forms the negative electrode. Instead of copper as in the wet cell, the positive electrode is a rod of carbon placed in the center of the cell. Small amounts of ammonium chloride, manganese dioxide and carbon, in paste form, are placed between the carbon rod and the zinc case.

The center terminal or binding post of the cell is the positive terminal, while the zinc case or outer terminal or binding post is the negative terminal. To connect two or more cells in *series* for your experiments, connect a bare wire from the positive terminal (center binding post of the battery) to the negative terminal (outside binding post) of another cell. This will leave one center terminal (positive) and one negative terminal for connections to your experimental apparatus, cells or solution. To add another cell, thus making three dry cells, in series, simply continue the inner terminal to the outer terminal connections. The more dry cells you connect in "series," the greater will be the voltage, or "electrical pressure," at the "end" or outer terminals.

EXPERIMENT 444—Making a Dry Cell Battery.

To make an experimental dry cell battery obtain an old, worn-out flashlight battery. Remove the contents by cracking the wax seal around the center post and then pulling the contents out of the battery case. Line the zinc case with a piece of blotting paper. Tie the black material (manganese dioxide and carbon), which you have removed from the original cell, in a small bag around the carbon rod. Now dip the bag and rod into a solution of ammonium chloride (2 teaspoonfuls in ½ tumbler of water) until it is thoroughly soaked. Place the damp bag in the zinc battery case and seal the top with paraffin, letting only the carbon rod project. Test this battery with your tangent galvanometer to see if it produces an electric current.

STORAGE BATTERIES

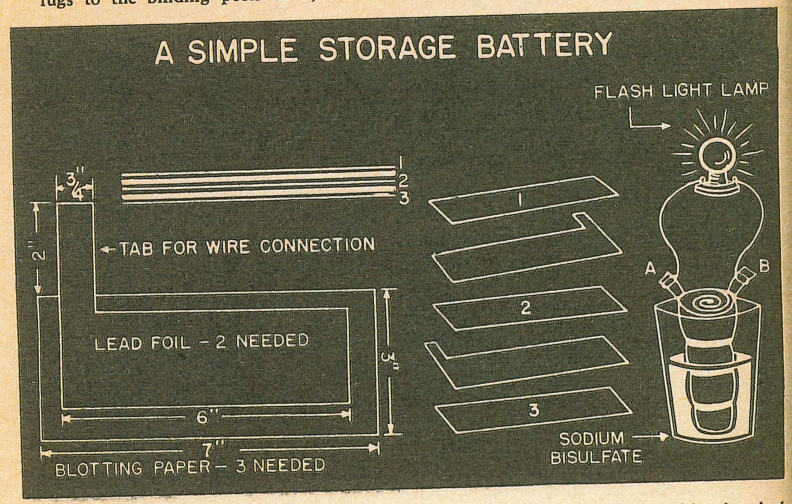
What a storage battery does in an automobile is well known. What it does not do is always attributed to the battery in a very matter-of-fact way. It does not store electricity. When the engine is running, the ammeter of the car shows charge—electricity is being applied to the terminals of the battery to charge it, not with electricity, but with chemical energy.

The storage battery consists essentially of a number of lead plates which form the positive and negative electrodes. These dip into a dilute solution of sulfuric acid and water. When the battery is receiving its charge, the positive plates become transformed into chocolate-colored lead peroxide while the negative plates are gradually changed into spongy lead. During the charging process, the electrolyte (sulfuric acid) becomes more concentrated and consequently the specific gravity of the solution is increased. This can be shown by testing the solution with a hydrometer before and after the battery is charged. When the starter button is pressed, the battery discharges and the acid, reacting with the plates, produces a current much as it was produced in your wet cell. It is possible to make a miniature storage battery which can be charged by connecting it to 2 or 3 dry cells that are connected in series. After disconnecting the dry cells, your storage battery can be used to light a flashlight bulb.

EXPERIMENT 445—Making a Simple Storage Cell.

Obtain 3 sheets of blotting paper measuring about 7 inches by 3 inches. Cut 2 sheets of lead foil as shown in the diagram, leaving a tab 2 inches long and ¾ inch wide on each sheet as shown at A and B. The lead foil can be obtained from pipe tobacco or from a package of tea. Do not use candy bar or chewing gum foil, as these substances are usually wrapped in aluminum or tinfoil. Lead foil is dull in appearance and does not crackle when bent. Starting with a piece of blotting paper followed by a piece of lead foil, place one upon another as shown in the diagram. Roll these into a cylinder, making sure that the lead tabs extend upward on the opposite sides of the roll. It may be advisable to bend the tabs slightly upward to insure a complete separation of the terminals. Make sure that the lead strips do not touch at any point. The purpose of the blotting paper is to separate or insulate the lead foil "plates" from each other.

In place of the sulfuric acid found in storage batteries, use a solution made by dissolving ½ teaspoonful of Sodium Bisulfate (No. 7) in a tumbler ¾ full of water. Lower the roll which you have formed into this solution, raising and lowering it several times to soak it thoroughly with the electrolyte. Connect the lead lugs to the binding posts of 2, or better 3, dry cells connected in series and pass



the current about 5 minutes. This charges the lead storage cell with chemical energy. Disconnect the cells and connect the storage cell to a flashlight bulb. The chemical energy which has been stored is now transformed into electrical energy, lighting the bulb. If you allow the cell to discharge through the bulb, its charge will gradually be reduced. However, it can be charged and discharged as often as you like.

EXPERIMENT 446—What Happens when Charging a Storage Cell.

Charge your storage cell by using 2 or 3 dry cells and then remove the roll from the solution of sodium bisulfate. Remove the rubber bands and unroll the blotting paper, thus exposing the lead plates. Notice that one of the plates is covered with a chocolate colored deposit. This is lead dioxide, also called lead peroxide, and was formed by the oxidizing action of the current on the plate when the battery was charged. When the battery is discharged, this lead peroxide reacts with the acid. It is the result of this chemical reaction, together with the one which takes place at the negative pole that produces the electric current. After dismantling the cell, wash your hands thoroughly to remove all traces of the sodium bisulfate solution.

ELECTROLYTES and NON-ELECTROLYTES

When sugar, glycerine, alcohol and most other organic compounds are dissolved in water, the molecules of the dissolved substance are simply mixed, unchanged, with the molecules of the water. Such a solution will not conduct an electric current and the dissolved compound is therefore called a non-electrolyte. When acids, bases and salts dissolve in water, unlike most organic substances, their solution conducts an electric current, and the dissolved substance is therefore called an electrolyte. The reason that such a solution conducts electricity is that it contains electrically charged particles, called ions. The metal in a solid crystal of a salt is in the form of positive ions, while the non-metallic part, such as chlorine, bromine, or the sulfate radical, is present as negatively charged ions. When the solid is placed in water, the ions dissolve and separate from each other so

that they are free to move as soon as the current is turned on. With bases, such as sodium hydroxide and calcium hydroxide, the same thing happens, but with acids the situation is a little different, since before they are dissolved they contain no ions. As soon as an acid is placed in water, however, a chemical reaction occurs and positive and negative ions are formed. All electrochemical reactions which involve the passage of an electric current through a solution or which bring about the production of electrical energy, are made possible by the presence of these ions in solution.

EXPERIMENT 447—Test for a Non-Electrolyte.

Dissolve several drops of glycerine in ¼ tumbler of water. Connect a clean, bare wire to one terminal of a dry cell, placing the other end of the wire in the tumbler of dilute glycerine. Connect another wire to the other terminal of the dry cell, and connect the free end to one of the wires of your tangent galvanometer. Put the end of the other wire of your galvanometer coil into the tumbler of dilute glycerine, but do not allow it to touch the wire already in the solution. There is no motion of the galvanometer needle, showing that current does not flow through a solution of glycerine in water. This is conclusive proof that glycerine is a non-electrolyte.

EXPERIMENT 448—Test for an Electrolyte.

Repeat the last experiment but instead of glycerine use several measures of Ammonium Chloride (No. 9) dissolved in ¼ tumbler of water. The needle will swing, showing that the ammonium chloride conducts an electric current when in solution. Ammonium chloride, therefore, is an electrolyte.

With the apparatus you have used in the preceding experiments, you can test any number of soluble substances to see whether they are electrolytes or non-electrolytes. It is necessary only to carefully clean the apparatus and introduce the solution to be tested into the tumbler.

The existence of ions in solutions of electrolytes and in melted salts makes possible some of our greatest electrochemical industries. Chlorine, hydrogen, oxygen, sodium hydroxide, metallic sodium, calcium, magnesium, and aluminum are examples of products made on a very large scale by passing an electric current through liquids containing ions. In previous sections of this book, we have done experiments to illustrate the production of hydrogen, oxygen and chlorine by electrolysis, that is, by passing a current through a solution of an electrolyte.

ELECTROPLATING

In electroplating, a solution of the salt of the metal which is to be deposited or plated on the object is used. To describe the action which takes place in electroplating, suppose that you are going to copper-plate an iron nail. First, make a solution of copper sulfate, then attach the nail to the zinc or negative pole of the battery or dry cell. Then attach a piece of copper to the carbon or positive pole of the battery. Now put the piece of copper and the iron nail which are attached to the battery in the copper sulfate solution and observe what takes place. The electric current may be considered as flowing from the positive to the negative pole, thus it would pass from the piece of copper, which is connected to the positive pole, through the copper sulfate solution, which contains positively charged copper ions and negatively charged sulfate ions. The positive copper ions are attached by the negative pole which is the iron nail, hence you have the copper deposited on the iron. At the positive pole copper ions are formed from the copper strip, thus bringing more copper into the solution to replace that which is deposited on the iron nail. Suppose you try this experiment and see what results you can obtain. You will get the best results by having the surfaces to be plated clean and bright.

EXPERIMENT 449—Copper Plating with an Electric Current.

Put 3 measures of Azurite (No. 39) and 8 measures of Sodium Bisulfate (No. 7) in a tumbler. Fill the tumbler half full of water and stir until the solids are all

dissolved. By means of a long piece of wire, connect a copper strip with the carbon or positive pole of your battery. Attach a bright iron nail or other bright iron object to another piece of wire, and attach this wire to the zinc or negative pole of your battery. It is a good idea to clean the metal with emery and then wash it with alcohol. Now put the copper strip and the nail or object which you wish to plate in the tumbler containing the solution and allow them to remain there for a few minutes. Be sure that they do not touch each other.

You will see that the piece of iron or steel is gradually turned red, due to a deposit of copper, while the copper strip is slowly dissolving and furnishing more copper to the solution. When the piece of iron has a sufficiently heavy coat of copper, remove it and examine the plating carefully.

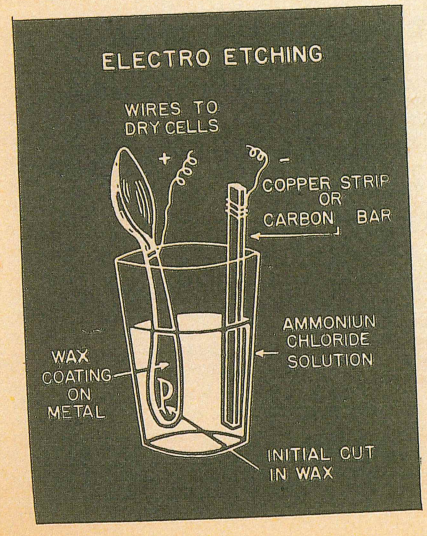
EXPERIMENT 450—Electroplating Your Name on Tools.

If you have a rubber stamp which you use to stamp your name in ink upon paper, you can use it to imprint your name upon tools or other metal objects by electrolysis.

Clean the ink from the rubber stamp with a cloth soaked in water. Dissolve as much candle wax as possible in ¼ test tube of carbon tetrachloride or naphtha. Clean the spot on the metallic object upon which you wish your name to appear. Pour the solution of wax on a small square of cloth, press your rubber stamp down upon this cloth as if it were an ink pad and imprint the stamp quickly on the object to be plated. The solvent will evaporate, leaving your name in an invisible film of wax upon the metal. Now place the object in a copper solution made as in the copper plating experiment. If the article to be plated is large, you can build a dam of wax or clay about the part to be plated and pour the copper sulfate solution into the tray thus formed. Connect the object to the negative pole of 2 or 3 dry cells connected in series and your copper strip to the positive pole, dipping the copper strip into the copper sulfate solution. The strip must not touch the object being plated. Plate for 15 minutes. Your name will appear standing out from a background of pure copper.

EXPERIMENT 450—Electroplating your Name on Tools.

You can perform the last experiment "in reverse," using wax to block out the portion upon which you do not desire to plate. To do this, spread a thin film of wax on the object to be plated, heat slightly to smooth the wax, then let it cool. Build a dam of clay or wax about the portion upon which you wish to plate and write your name with a needle or scriber through the wax film. Pour some of your copper sulfate solution into the dam, connect the tool to the negative side of the two dry cells in series, and hold a wire connected with the positive side of the dry cells in the copper sulfate solution. Plate for 10 minutes. Wipe the article thoroughly. Your name will be plated in copper upon the tool.



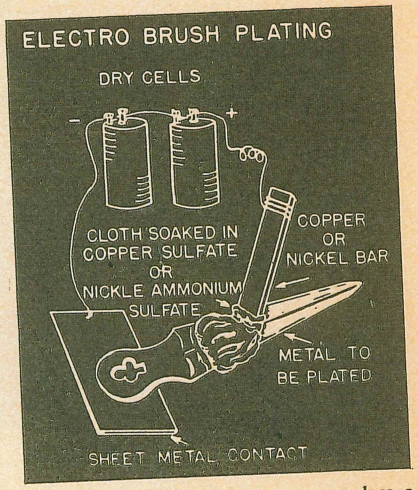
EXPERIMENT 452 — Electro-etching.

You can etch or engrave your name, initials, or any drawn design upon metallic objects. Coat the object to be engraved with wax as in the preceding experiment, and scribe your design through the wax. Connect the object to the positive pole of three dry cells connected in series and place in it a glass containing several measures of Ammonium Chloride (No. 9) dissolved in ½ glass of water. Connect a copper strip or carbon rod to the negative pole of the dry cells. During electrolysis the chlorine ions are converted to atoms of chlorine gas at the positive electrode, which attack the metal object on the portions unprotected by wax.

Sometimes the object which one wishes to electroplate is so large that it would be necessary to build a special electroplating bath in order to plate sucessfully. The electrochemist therefore has devised a method of plating in which no bath is necessary. This method is known as electroplating by "brushing". In the following experiment you can use this method to plate some object which is not easily introduced into a bath.

EXPERIMENT 453—Plating with the "Electroplating Brush."

Connect a wire leading from the negative terminal of two or three dry cells connected in series to a piece of iron, such as the blade of a knife, or any other iron object upon which you would like to plate a coat of copper by the brush process. Be sure that the iron is clean and bright. Wrap one end of a copper strip or copper bar with absorbent cotton or soft cloth so that a thick pad is formed. Connect the copper bar to the positive terminal of your dry cell series by means of a copper wire. Make a solution of copper sulfate by dissolving about 5 measures of Azurite (No. 39) in a small test tube of water containing 10 measures of Sodium Bisulfate (No. 7). Heat slightly to cause all the azurite to dissolve. Dip the cotton-padded end of the electroplating brush which you have made into the copper sulfate solution to moisten it thoroughly, and rub it across the iron object which you wish to copper plate. You will see the copper deposited upon the iron as the stroke is made. By repeating the stroking process, the copper film can be made thicker and thicker. The cloth pad buffs the metal slightly, at the same time that it plates, thus insuring a bright surface of copper on the iron. If the plating is too dark in color and tends to rub off, then use only two dry cells as the source of current, or place a resistance in series with one of the wires from the cells. Instead of wrapping the absorbent cotton about your copper bar or strip, you can wrap it about a flashlight battery carbon and use this as the brush, but you will have to dip the cotton pad into your copper sulfate solution more frequently to replenish the copper which was lost by plating out on the iron.



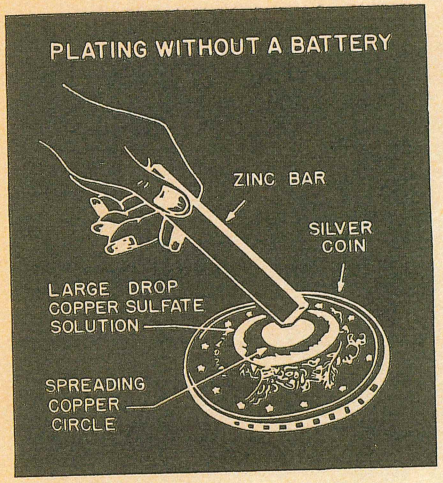
Plaster vases, small clay vessels, or objects molded from similar materials do not conduct an electric current. However, it is possible to plate these objects by first brushing them with powdered graphite so that a uniform black film is formed on their surface. This film of graphite will conduct the current and make electroplating possible. After brushing with graphite the object is connected to a source of current and placed in a plating bath as before. Sometimes the object to be plated is dampened with a solution of silver nitrate, then exposed to hydrogen sulfide gas which precipitates silver sulfide on the surface of the object. The wire is then attached and the plating carried out in the usual manner.

EXPERIMENT 454—Plating a Clay or Plaster of Paris Object.

Thoroughly cover the object which you wish to plate with a coating of graphite by rubbing powdered graphite on the object until it has a shiny black coating. Now, by means of a wire, connect it to the negative terminal of a dry cell and place in the copper sulfate solution which you used for plating. With your copper strip connected to the other (positive) pole of the cell and placed also in the copper sulfate solution, the work can be plated with copper.

In most electroplating operations, a source of direct electric current is utilized to bring about a deposit of metal from solution. For plating small objects, however, one can sometimes use the difference in solution potential (tendency for a metal to dissolve) of two metals to bring about the desired result without any use being made of electric current from a battery.

EXPERIMENT 455 — Plating a Silver Coin Without a Battery.



Dissolve two measures of Azurite (No. 39) in ¼ test tube of water containing 4 measures of Sodium Bisulfate (No. 7). This will form a solution of copper sulfate. Place a large drop of this solution upon a silver coin which has been thoroughly cleaned with emery cloth or steel wool. Clean one corner of your zinc strip and thrust it down into the drop of copper sulfate solution on the coin, but do not touch the coin with the zinc. Observe that the zinc which is wetted by the copper sulfate solution becomes plated with copper. The copper is plated simply by displacement, the zinc having a greater solution potential, or preferring to go into solution more readily than the copper metal. Now thrust the zinc bar down through the drop of copper sulfate solution so that it touches the silver coin. You will immediately notice a film of copper being plated out upon the silver coin, spreading in circular form until the whole silver coin in contact with the liquid drop is plated.

EXPERIMENT 456—Copper Plating by Immersion.

Dissolve 4 measures of Sodium Bisulfate (No. 7) in a test tube ½ full of water. Now add to this solution 1 measure of Azurite (No. 39). Warm the solution gently, to completely dissolve the azurite. Now dip a clean piece of steel into the solution and notice that almost immediately a red coating is formed on the steel. Remove the steel from the solution after half a minute or so, rinse it off and dry it with a soft cloth. The steel will be copper plated.

EXPERIMENT 457—Tin Plating.

Place 6 measures of Sodium Bisulfate (No. 7) in a tin cup and fill about ¾ full of water. The cup used should be bright and must be tinware (that is, sheet iron coated with tin). In this experiment, part of the tin from the cup is dissolved and re-deposited on the article to be plated.

Now drop a bright clean penny into the cup, set it on the stove and boil the solution until the water is nearly all driven off. Watch the penny; it will gradually be covered with a bright, silvery plating of tin. The penny used must be bright and clean, as a tarnished penny will not plate well. A dirty penny may be cleaned by boiling it in strong vinegar for 10 to 15 minutes. Cleaning can be hastened by adding 3 measures of salt.

Certain metallic chemical elements possess the property of being able to displace certain other metals from a solution of their salts. This property is known as the relative electro-positiveness of metals and is made use of in a great many ways. This is fundamentally the same sort of reaction which occurs in electroplating. The difference is that here the two metals and the solution form their own battery and produce a small current. The experiments on copper, and tin plating without the use of a battery made use of this principle.

EXPERIMENT 458—Electro-positive Action of Silver.

Dissolve 2 measures of Sodium Bisulfate (No. 7) in a test tube ¾ full of water. Pour this solution into a clean tumbler and put a silver coin in it. Notice that there is no action, as the acid does not dissolve silver at all. Now put a zinc strip in the solution, arranging it so that it will touch the silver coin. You will find that almost at once bubbles of gas begin to come off from the coin while bubbles collect on the strip of zinc.

Experiment 459—Electro-positive Action of Copper.

Repeat the preceding experiment, using a copper cent instead of a silver coin. Notice that the same action takes place when the zinc is introduced into the solution and touches the copper.

EXPERIMENT 460—Electro-positive Action of Iron.

Try this experiment again, using an iron nail. In this case, too, you will find that there is no rapid action until the zinc is introduced into the solution so that it touches the iron.

DIRECT AND ALTERNATING CURRENT

Everyone knows that there are two types of electric current, direct and alternating, but few people ever stop to analyze the significance of these names. Direct current is electrical energy that flows continuously along a wire in one direction. Alternating current flows along a wire in one direction, rises to a maximum value, then drops to zero and begins flowing in the opposite direction. In 60 cycle current, the current changes its direction of flow 120 times every second.

If an attempt is made to use alternating current in electroplating, the objects will not be plated, because although metal is deposited when the flow is in one direction, when the flow is reversed the metal is immediately removed.

In order to determine whether current is alternating or direct, advantage may be taken of the fact that if a current is allowed to electrolyze a solution of salt water, sodium hydroxide is formed about the negative pole. By using an indicator, such as phenolphthalein, in the salt solution, one can tell by the pink color formed by the basic action of sodium hydroxide which pole is negative, and if the current is alternating, a pink color will form about both poles.

EXPERIMENT 461—Making Pole Testing Paper.

Wet a strip of white paper with a solution of ordinary table salt, about 5 measures dissolved in a half test tube of water. Add a drop of Phenolphthalein Solution (No. 22) to the wet paper. Place the wires connected to the current source upon the paper and about ½ inch apart. If a pink spot forms in several seconds about one wire, the current is direct and that particular wire is the negative wire. If the paper about both wires turns pink, the current is alternating current. If no coloration results, the implication is that no current is flowing or that the current is too weak to cause the electrolysis of the salt water in the paper.

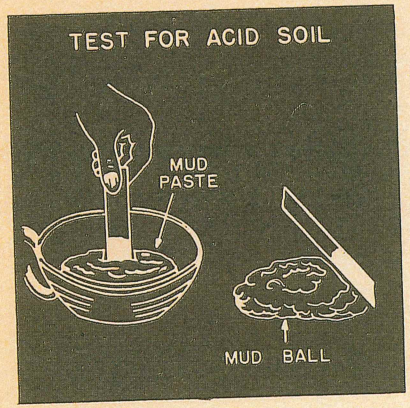
This sort of pole testing paper may be dried and kept on hand for instant use. Simply moisten it with water and touch to it the wires whose polarity you wish to test.

CHEMISTRY IN AGRICULTURE

Today farmers are vitally interested in the application of chemistry to their farming methods. The farmer knows that by using certain types of fertilizers, certain fungicides, and certain insecticides, he stands a much better chance of obtaining a satisfactory yield from his crops. By obtaining disease-free seed, and knowing the chemical constituents of his soil, it

is possible for him to be certain (weather permitting) of reaping a better return from his land.

The chemical composition of soil varies greatly in different localities depending upon the type of rock from which the soil was derived. All soils contain silicon and aluminum, so plants never suffer for lack of them, as they sometimes do when there is insufficient nitrogen, phosphorus or potassium. Many soils contain these three vital elements, but they are frequently combined so that they cannot readily be absorbed by the plants. This may sometimes be remedied by treating the soil with an alkali, such as hydrated lime (calcium hydroxide).



EXPERIMENT 462—Test for Acid Soil.

Obtain a sample of soil and mix with it sufficient clear water to make a thin paste. Imbed a strip of blue Litmus Paper (No. 18) in the mixture and allow it to remain for half an hour. Withdraw carefully and rinse in clear water. If the paper has turned pink, the soil is acid.

This test may also be carried out by making a mud ball, pressing a piece of blue litmus paper against it, and allowing it to stand half an hour.

Sorrel and moss thrive in acid soil. Clover, alfalfa and peas will not grow, and wheat, corn and most crops grow poorly in acid soil.

EXPERIMENT 463—How to Test for Alkalinity.

Secure a sample of soil from a different place from that used in the preceding experiment and mix it with water to obtain a thick paste. Inbed a strip of red Litmus Paper (No. 18) in the mixture. Allow the paper to remain in this position for ½ hour after which carefully withdraw it and rinse it in running water. If the paper has turned blue, the soil which you tested is alkaline.

Plants take carbon dioxide from the air and combine it with water from the soil to form cellulose, starch and sugar. This process goes on in their green leaves, using energy obtained from sunlight. If extra quantities of carbon dioxide are supplied to a plant, it will grow faster. In other words, carbon dioxide acts as a sort of fertilizer.

EXPERIMENT 464—The Effect of Carbon Dioxide on Plant Growth.

Obtain two 1-pint fruit jars and put about 2 inches of damp earth in each. Plant 3 or 4 beans in each jar, being careful to plant all the beans at the same depth. Now generate carbon dioxide gas by adding vinegar to baking soda. Generate the gas in a large milk bottle and carefully pour it into one of the fruit jars, being careful not to allow any of the liquid to get into the jar. Label and seal both jars tightly. One contains carbon dioxide and the other does not. Set in a quiet place and after the beans have sprouted, you will find that the plants in the jar containing carbon dioxide grow faster than those in the other jar.

FERTILIZERS

Fertilizers are used on soil to replace the natural elements which have been consumed by the growth of crops. Nitrogen, phosphorus, and potassium are the elements which most frequently require replacement.

Nitrogen may be replaced in soil by the use of nitrates, ammonium salts, refuse from slaughter houses (tankage), ground fish meal, cottonseed meal, and cotton oil pomace.

Phosphorus is derived from fertilizers made from raw or steamed bones, bone tankage, or mineral phosphates.

Potassium is replaced in the soil almost wholly by the use of the chemical compounds, potassium sulfate and potassium chloride. Wood ashes formerly constituted a large portion of the potassium fertilizers.

EXPERIMENT 465—How to Make a Garden Fertilizer.

An excellent garden fertilizer can be obtained by mixing 2 parts of wood ash with 1 part of poultry manure. This fertilizer will be rich in potassium, phosphorus and nitrogen.

Place about 1 tablespoonful of this fertilizer in the soil around the roots of some potato plants in your garden. Notice the rapid growth, the bright green color of the leaves, and the early maturity of the fruit resulting from the proper use of this fertilizer.

Some interesting recent experiments have shown that plants can be grown at a rapid rate and to a large size by supplying them with nutrients in solution. Thus crops can be raised without any soil, the roots of the plants simply being dipped into a solution to which is added the substances needed for growth.

Other experiments have shown that the presence of very small amounts of certain organic compounds, called plant hormones, enormously increases the rate of growth of various plants, and by special treatments, it has been possible to produce certain fruits, such as tomatoes, without seeds. Furthermore, it has been found that a powerful drug, called colchicine, when properly applied, is able to produce deep-seated changes in plants. Entirely new varieties have arisen due to its action upon certain seeds.

INSECTICIDES AND FUNGICIDES

To protect plants against the ravages of insects, huge tonnages of insecticides are used every year. The most important are perhaps certain compounds of lead and arsenic, such as lead arsenate and calcium arsenate. These substances are either suspended in water and used as a spray or they are spread over the plants in the form of dust. Insects feeding on the leaves or fruit eat the poisonous insecticides and are killed.

Sucking insects, whose mouth-parts penetrate leaves and twigs and suck the sap of the plant, are not harmed by these arsenical insecticides. For such insects it is necessary to apply substances which clog their breathing pores, which give off poisonous fumes, or which penetrate the body of the insect without being eaten. Various oil emulsions and nicotine preparations accomplish this.

A serious disadvantage of the arsenic insecticides is that they are poisonous to man. The removal of spray residues from apples and other fruits has become a serious problem, and in spite of the efforts of the Department of Agriculture, it is fairly certain that some people are getting more lead and arsenic than is good for them. A possible remedy is the wider use of vegetable insecticides, such as the compound rotenone extracted from derris and other roots. These substances are poisonous to insects, but not to the higher animals.

Fungicides are preparations used to kill fungi, tiny plants which infest and injure some of our most important crop and ornamental plants. The task of a fungicide is to kill the fungus without injuring the plant on which it is growing. The most familiar are compounds of sulfur and of copper with lime.

EXPERIMENT 466—Lime-Sulfur Solution.

Place in a test tube 2 measures of Calcium Oxide (No. 20) and 1 measure of Sulfur (No. 1). Fill the tube half full of water and boil the mixture over the flame of your alcohol lamp until the solution assumes a yellow color.

Add the yellow solution to a glass ¼ full of water and you will have a lime-sulfur insecticide which, when sprayed on fruit trees, will destroy fungus growths and also certain insects.

EXPERIMENT 467—How to Make an Oil Emulsion Insecticide.

Prepare some of the lime-sulfur concentrate as you did in the preceding experiment. Instead of adding to water, however, this time pour it into about 4 teaspoonfuls of very thin motor oil in a small pan. Add a piece of kitchen soap approximately ½" square. Boil this mixture until the soap, oil and lime-sulfur form a cream or emulsion. The soap must be added in order to emulsify the oil with the water solution of lime-sulfur. Under ordinary conditions, water and oil will not mix. Add slowly to this emulsion ½ pint of water with constant stirring, and you will have a very effective spray material for exterminating "sucking" insects.

EXPERIMENT 468—Using Sulfur As a Fungicide.

Many times gardeners are troubled with fungi which produce injurious effects upon plants. Mildew is a fungus which can easily be destroyed with sulfur. Dust a little Sulfur (No. 1) on the soil around the plant and the mildew will be destroyed. If one application does not suffice, repeat the procedure. A flat one-inch paint brush is a good means of dusting sulfur on the soil around plants.

EXPERIMENT 469—Making a Copper Fungicide.

Bordeaux mixture is an effective fungicide for many purposes. It derived its name from a district in France where the favorable effects of lime and copper salts were first used to destroy harmful fungi on grapevines. It is essentially an insoluble mixture of compounds of copper and calcium. These compounds are formed as a finely divided precipitate when a solution of copper sulfate is added to lime water.

To ½ test tube of water, add ½ measure of Azurite (No. 39) and 1 measure of Sodium Bisulfate (No. 7). Boil this mixture gently over the flame of your alcohol lamp until the substances dissolve.

In another test tube place 4 measures of Calcium Oxide (No. 20), add ½ test tube of water and boil it for a moment.

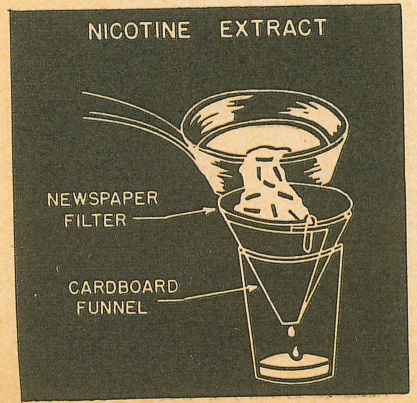
Allow both test tubes of liquid to cool and then pour the contents of each into a small cup with constant stirring. At once you will notice the formation of a precipitate composed of extremely fine particles. These suspended particles are the active part of Bordeaux mixture.

EXPERIMENT 470—Testing Bordeaux Mixture for Free Copper.

Unless an excess of lime is present in your Bordeaux mixture, the preparation may contain sufficient soluble copper to be injurious to the foliage of plants.

To test for soluble copper, filter a little of your mixture. To the filtrate thus obtained add 3 or 4 drops of vinegar. To a test tube ¼ full of water, add ½ measure of Sodium Ferrocyanide (No. 6). Now add 2 drops of this solution to the filtrate from the mixture. If soluble copper is present, a dark reddish-brown color will appear due to the formation of copper ferrocyanide.

When Bordeaux mixture is applied to plant foliage, the copper compounds gradually dissolve due to the action of the carbon dioxide given off by the leaf. This soluble copper is poisonous enough to kill most fungi, without injuring the plant itself.



EXPERIMENT 471—Nicotine Extract

Place the tobacco from a cigarette in a small pan containing half a glass of water. Allow this to boil for 5 or 10 minutes, and then filter. Now dissolve 2 measures of soap in ⅓ tumbler of water and add the nicotine extract. This mixture is often used for spraying house plants infested with aphids.

HOUSEHOLD CHEMISTRY

The importance of chemistry in our everyday life about the home is being more fully realized, and many schools are now offering special courses in this subject. The following experiments indicate the large field which is covered by household chemistry.

EXPERIMENT 472—The Use and Care of Aluminum Ware.

Obtain an old aluminum pan and put some vinegar in it. Warm slightly and notice the corrosive action which the vinegar has on aluminum. Add a small quantity of salt to the vinegar and notice that the corrosive action is increased. Put a small quantity of lye in the aluminum pan, dissolve it in water, and notice that this corrodes the aluminum.

This experiment shows that acids and alkalies should not be put on aluminum ware for cleaning or polishing. Neither should acid foods be prepared or kept in aluminum vessels. Salad dressing containing vinegar and salt, for example will corrode aluminum pans very badly and also introduce aluminum compounds into food.

EXPERIMENT 473—Cooking in Copper Vessels.

Coil a short piece of bare copper wire into a spiral and place it in a cup containing a little strong vinegar. Let the wire remain partly immersed in the vinegar for a day or two. In the presence of oxygen of the air, the copper is oxidized and reacts with the acetic acid of the vinegar to form blue copper acetate. Pour a little of the copper acetate solution into a test tube and add a few drops of household ammonia. A deep blue color will be formed, showing that the solution contains copper. You find, therefore, that copper vessels are not good for cooking acid foods, as the copper will oxidize and dissolve sufficiently to contaminate the food with poisonous copper compounds.

EXPERIMENT 474—Why Zinc Vessels Are Not Used for Food.

Test the solubility of zinc in acids, for example, warm vinegar. Also test the solubility of zinc in alkalies, such as a strong lye solution. We find that zinc is attacked by acids and alkalies. Therefore, foods should not be kept or prepared in zinc vessels. Sodium chloride solution also attacks zinc and as the soluble compounds of zinc are poisonous, it would be dangerous to use zinc vessels for preparing or storing foods. Galvanized vessels are made of iron, coated with zinc. Therefore, the same principles apply to galvanized vessels as to zinc vessels.

EXPERIMENT 475—Testing Silver Polish for Ammonia.

Silver polishes very frequently contain ammonia or ammonium salts and an alkali, as this aids in cleaning silver. Put a small quantity of the silver polish which you wish to test, possibly 1 or 2 measures, in a test tube. Add 1 measure of Calcium Oxide (No. 20) and 2 or 3 drops of water. Heat the mixture for a few moments and smell cautiously from time to time, first removing the tube from the flame. If ammonia is present in the polish, you will be able to smell it at the mouth of the tube.

EXPERIMENT 476—Cleaning Silver with Sodium Thiosulfate.

We have learned that sodium thiosulfate diccolves silver sulfide. Prepare a solution of 2 measures of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. Moisten a clean soft cloth in the solution and rub a piece of silverware which is tarnished with black silver sulfide. You will find the silver sulfide comes off very easily.

EXPERIMENT 477—Cleaning Silver by the Electrolytic Method.

Pieces of silverware, either plated or solid, which are tarnished can be cleaned so that they look like new without rubbing with silver polish as we learned in a previous experiment. Obtain an old aluminum pan large enough to hold the silver you wish to clean. Fill the pan with hot water. Add 1 teaspoonful of sodium chloride (common table salt) for each quart of water used. Now put in the pieces of silver which you wish to clean so that they are completely covered with water. Leave them there about one-half minute, longer if the silver is very dirty. Remove the silver, rinse in warm water and dry with a soft cloth. This method is less harmful to silver than the usual polish which wears away a little of the silver.

The aluminum is corroded while the silver is cleaned because it replaces the silver in the compound which forms the tarnish, so it is best not to leave the silver in the pan any longer than necessary. If an old leaky aluminum pan is available, this may be used by setting it in a large enameled or agate ware pan.

EXPERIMENT 478—Corrosive Action of Alkalies on Glass.

Obtain some strong lye and dissolve about a teaspoonful of it in a glass ¾ full of water. Allow this solution to stand for a day or two and then pour it out. Examine the glass and notice how the alkali has etched or corroded the inside of the glass. Lye is sodium hydroxide or caustic soda.

EXPERIMENT 479—Testing Laundry Bluing for Ultramarine.

Bluing is used in washing to neutralize the yellow tinge of clothes and to make them appear whiter. Put ½ measure of blue in a test tube ¼ full of water. Add 3 measures of Sodium Bisulfate (No. 7) and shake the test tube thoroughly. If the blue is ultramarine it will be turned colorless by the sodium bisulfate and hydrogen sulfide gas will be given off which you can smell at the mouth of the tube.

EXPERIMENT 480—Testing Laundry Bluing for Prussian Blue.

Prussian blue is very commonly used for laundry bluing. To test for this substance, place ½ measure of blue in a test tube ¼ full of water. Add 2 measures of Sodium Carbonate (No. 4) and warm gently. If the solution changes from blue to brown or dark red, the substance is Prussian blue.

EXPERIMENT 481—Testing Laundry Bluing for Indigo.

Indigo is frequently used for laundry bluing. To test for indigo put 1 measure of blue in a clean test tube and heat it. If a purple vapor comes off and a very unpleasant odor is noticed at the mouth of the tube, the substance is indigo.

If the blue which you were testing does not correspond to any of the above tests, it is probably some other type of organic dye.

EXPERIMENT 482—Testing Face Powder for Boric Acid.

Mix 2 measures of the face powder in a cup with 2 measures of Sodium Bisulfate (No. 7). Transfer the mixture to a spoon, add a few drops of alcohol, and set on fire. If the alcohol burns with a green flame, the face powder contains boric acid or borax. The same test may be applied to other cosmetic and toilet preparations.

EXPERIMENT 483—Testing Face Powder for Starch.

Add 1 measure of face powder to a test tube ¼ full of water, shake thoroughly and then boil for a minute or two. Set the tube aside for 10 or 15 minutes, until the liquid is perfectly cool and then add 1 drop of tincture of iodine. A blue color indicates the presence of starch. Starch is somewhat undesirable in face powder since the skin of some persons is sensitive to it. You may also try this test on toothpastes, tooth powders and other cosmetics.

EXPERIMENT 484—Testing Depilatories for Sulfides.

Proprietary hair removers sometimes contain barium, calcium or sodium sulfides. Place a small quantity of the preparation in a test tube with 2 measures of Sodium Bisulfate (No.7) and a few drops of water. Warm gently over a flame and then remove the test tube from the flame and smell at its mouth for the odor of hydrogen sulfide. Test also with a moistened piece of Sulfide Test Paper (No. 31). Soluble sulfides have a solvent action on hair, but unfortunately, the chemical composition of the skin is very similar to that of the hair, so that if they are allowed to act for too long, they may injure the skin as well as remove the hair.

Recently, depilatories containing derivatives of an organic compound known as thioglycollic acid have been put on the market. They are said to irritate the skin less than the older type.

EXPERIMENT 485—Testing Silver Plate.

To find out whether an article is plated with nickel or with silver, prepare a solution of 4 measures of Sodium Bisulfate (No. 7) in a test tube half full of water. Put the article which you wish to test in this solution and warm it for a few minutes. If a greenish solution is formed, the article contains nickel. Test the solution by adding Sodium Carbonate (No. 4) until it no longer effervesces and then generate hydrogen sulfide gas, and pass the gas through your delivery tube into the solution. A black precipitate indicates nickel.

EXPERIMENT 486—How to Make Windshield Compound.

Sometimes when driving a car in rain, snow, sleet, fog or frost, it is quite difficult to keep the windshield so that you can see through it. A simple and effective compound can be made of soap and glycerine which will prevent the windshield from becoming fogged or streaked with rain or ice. The same compound can be used on eye glasses, and windows to keep them clear in bad weather.

Shave some white soap by scraping with a knife. Put 25 measures of these soap shavings and 4 or 5 drops of glycerine in a cup. Add a few drops of water and stir thoroughly with a stirring rod or spoon until a smooth stiff paste is formed. Keep this paste in a tight container, such as a screw cap jar, and when needed, take a little on your fingers and rub over the outside surface of the glass to be kept clear.

EXPERIMENT 487—Testing Battery Terminal Corrosion for Copper.

Dissolve 2 or 3 measures of the corrosion in a test tube ½ full of water. Pour a drop of this solution into a test tube containing 1 measure of Sodium Carbonate (No. 4) dissolved in a test tube ¼ full of water. If a greenish precipitate appears, it would indicate that the corrosion contains copper.

Test a small piece of the corrosion in the alcohol lamp flame by putting a little on the end of the nichrome wire. Do you get a green copper color in the flame? Dip a clean iron nail into a solution of the corrosion and after a minute or so withdraw it; see if there is a thin coating of copper left on the nail.

EXPERIMENT 488—Testing Battery Terminal Corrosion for Iron.

Add a few drops of the solution of the corrosion to a test tube containing a little Sodium Ferrocyanide (No. 6) dissolved in water. A dark brown precipitate indicates the corrosion contains copper, or a blue precipitate or blue color indicates the presence of iron.

EXPERIMENT 489—Testing Storage Battery Terminal Corrosion for Sulfate.

Since sulfuric acid is used in automobile storage batteries, the corrosion must contain sulfates. Prepare a solution of 2 measures of Strontium Chloride (No. 10) in a test tube ¼ full of water. Add to this a few drops of a solution of the storage battery corrosion which you have been testing. The formation of a white precipitate of strontium sulfate proves the presence of sulfate.

EXPERIMENT 490—Testing for Acid Mouth.

Cut a piece of blue Litmus Paper (No. 18) and place it in your mouth for a moment. If when you remove the paper it is turned red, you have an "acid mouth."

Decayed teeth or an upset stomach are generally the cause of "acid mouth." Perhaps you have seen advertisements of tooth paste which offer test papers for "acid mouth." These test papers are blue litmus paper.

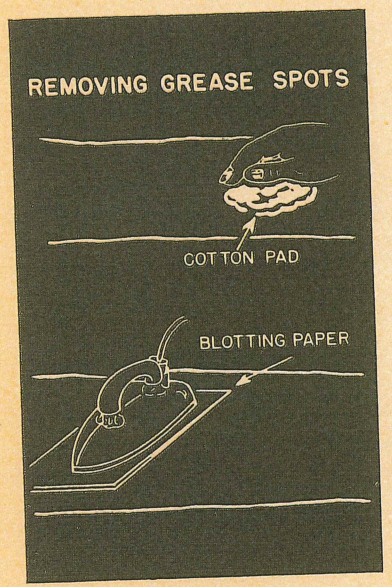
EXPERIMENT 491—Testing Tooth Powder for Ammonium Compounds.

Some tooth powders now contain ammonium compounds which are said to prevent tooth decay. Put 2 measures of tooth powder in a test tube and add 2 measures of Calcium Oxide (No. 20) and 2 or 3 drops of water. Warm this mixture gently and occasionally smell at the mouth of the test tube and see if you can detect the odor of ammonia. Such tooth powders usually contain an ammonium phosphate and sometimes also urea, which is an ammonia derivative of carbonic acid.

Many stains on fabrics can be quickly and economically removed at home by making use of some elementary knowledge of chemistry.

EXPERIMENT 492—How to Remove Iodine Stains.

The brown stain of iodine itself or the blue color, which it forms in contact with starch, can be removed from flesh or from fabrics by means of Sodium Thiosulfate (No. 16). Dissolve a little of the thiosulfate in water and treat the stain with it repeatedly until it slowly disappears.



EXPERIMENT 493—Removal of Grease Spots.

If warm water and soap cannot be used because of the nature of the fabric, a substance to absorb the grease should be applied. Place a piece of clean blotting paper on both sides of the fabric and apply gentle pressure with a warm electric iron to the upper piece of blotting paper. The grease softened by the heat is gradually taken up by the absorbent paper. Any remaining grease can probably be removed by laying the fabric on another piece of clean blotting paper and rubbing the spot with a soft cloth or wad of cotton which has been wet with carbon tetrachloride or naphtha. Rub from the edge of the spot toward the center. The solvent is absorbed into the blotting paper underneath the spot and carries the grease with it.

EXPERIMENT 494—Removing Stains with Lemon Juice.

Stains caused by ink or iron rust can often be removed with lemon juice. If the stain does not yield at once, moisten it well with the lemon juice and then place it in the bright sunlight for a few minutes.

EXPERIMENT 495—Bleaching Powder as a Stain Remover.

A great many fresh stains are best removed by pouring boiling water through the cloth. If the stain is not completely removed in this way, a bleaching agent may be tried. Stains due to ordinary and colored inks, to fruits and vegetables, to grass, and to tea and coffee may be treated as follows. Shake 1 measure of Calcium Hypochlorite (No. 38) in a test tube half full of water and soak the stain in this liquid for 5 minutes. Then transfer it to some lemon juice which has been squeezed into a cup. If the stain persists, repeat the treatments.

After the stain is removed, it is best to destroy the chlorine remaining in the cloth by soaking it for a few minutes in a solution of 2 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water, first rinsing the material thoroughly in water. Bleaching powder must not be used on wool or silk since it weakens or destroys them. It should not be used on colored cloth unless a test has shown that the color is not bleached by it.

THE CHEMISTRY OF FOODS

The food which we eat consists chiefly of compounds or mixtures of compounds containing the elements carbon, hydrogen, oxygen and nitrogen. There are three principal classes of foods—carbohydrates, fats and proteins. Carbohydrates include sugars and starches and are found chiefly in vegetable foods, while fats and proteins are found in animal foods as well as in plants, especially in their seeds.

It is most important that the food we eat be clean and pure and furthermore, that it agrees with us, as no doubt you have learned from experience. The whole process of digestion and absorbing food into our bodies is but a series of perfectly balanced chemical reactions and if we eat something that causes these reactions to get out of balance, a case of indigestion usually results. The chemist knows that certain foods are hard to digest because they do not dissolve easily in the digestive juices. Also, he can tell that certain substances used to flavor, color or preserve food are harmful, while others are perfectly safe to use.

The Federal government and various local health boards maintain chemical and bacteriological laboratories in many of our large cities for the purpose of testing and preventing the sale of foods which contain harmful ingredients or which have deteriorated because of improper storage.

FATS AND FATTY ACIDS

When we speak of fats, we generally have in mind substances like pork fat (lard), chicken fat, or butter. However, liquids such as cod liver oil, olive oil, and linseed oil, if they are cooled to low enough temperatures, will become solid just like any other fat. These oils are much like the solid fats in chemical nature and so they are called fatty oils. They are simply fats which happen to melt below ordinary temperatures. Butter, which is ordinarily a solid fat, becomes an oil on very hot days.

Of course there are other kinds of oils beside fatty oils. Mineral oils, for example, which come from petroleum, and the essential oils used in perfumes and flavoring extracts, are quite different in chemical nature from the fats.

Everyone knows that ice melts at zero degrees Centigrade (32 degrees Fahrenheit), and that water freezes at exactly the same temperature. In other words, the melting point of solid water and the freezing point of liquid water are the same. This is true of most substances, for example, sulfur and lead; but with fats, the freezing or solidifying point may be many degrees below the melting point.

EXPERIMENT 496—The Melting and Freezing Points of Fats.

Warm a lump of butter very gently in the spoon until it just melts. Feel the bottom of the spoon to note its temperature. If you have warmed it carefully, the the butter will melt before the bowl of the spoon becomes too hot to hold in your hand. Now let it cool, and you will find that the butter remains liquid at much lower temperature. You may even have to put it in cold water to make it solidify.

Certain liquids such as ether, chloroform, carbon tetrachloride and naphtha are known as fat solvents, since fats will dissolve in them very readily.

EXPERIMENT 497—Dissolving Fats.

Place a piece of butter about the size of a pea in a test tube and then fill the tube ¼ full of carbon tetrachloride or naphtha. Close the mouth of the tube with

your thumb and shake until the butter is dissolved. The solution may be a little cloudy because butter contains, in addition to fat, a small amount of protein from the milk, a little water and generally some salt.

Fill a cup nearly full of water and very gently pour a few drops of the butter solution onto the surface of the water. It will remain on top of the water in a thin film. After the solvent has evaporated, look across the surface of the water toward the light and you will see a film of fat remaining on the water.

Fats will dissolve in alcohol when it is heated. Try dissolving a little fat in rubbing alcohol. Use only a spoonful of alcohol and heat it cautiously in a test tube with a small portion of fat. Keep your face at a distance and be very careful not to spill hot alcohol as it takes fire readily. When the fat is dissolved, cool the test tube and watch the crystals of fat form in the tube.

In laboratories where the composition of foods is determined, fat solvents are used to extract the fat from food samples. The solvent is then evaporated off, and the residue is weighed on a delicate balance.

EXPERIMENT 498—Testing for Fat in Foods.

Practically all foods contain at least traces of fat, and many contain enough to be detected easily by the following simple method. Try this experiment on the kernel of any kind of nut, on an olive, a piece of cheese or a piece of fish.

Crush the sample in a cup with the back of a spoon. Put about 4 measures of the crushed sample in a test tube and fill the tube ¼ full of carbon tetrachloride or naphtha. Close the mouth of the tube with your thumb and shake it vigorously for several minutes. Now place a small wad of cotton in the mouth of the test tube and pour a few drops of the solvent from the test tube onto the surface of some water in a cup. The cotton will prevent the fibers of the sample from pouring out with the clear solvent. The drops of solvent containing the fat will float on the water in the cup. If you give the solvent time to evaporate and then look carefully across the surface of the water toward a light, you will see a grease film showing that the food contained fat.

EXPERIMENT 499—Showing that Rancid Fat Contains Acid.

Every cook knows from experience that if fats are left in a warm place they will become rancid and unfit for use. The change is due to the chemical action of the oxygen and moisture of the air, which breaks the fat down into a variety of ill-smelling compounds. A certain amount of weak acid is also formed.

Set a piece of butter or other fat in a warm, light place, and press a moistened piece of blue Litmus Paper (No. 18) against it. Examine it occasionally during a week or two, and notice the odor which develops. Enough acid is often formed to turn the litmus paper red. If the fat is butter, you will notice that the yellow color disappears. It is this yellow coloring matter of butter which is converted into vitamin A when we eat it. Hence, the vitamin value of butter is destroyed when it becomes rancid. Artificial coloring matter is often added to butter, and of course when this is done, the color of the butter is no indication of how much vitamin value it has. Artificially colored butter may not lose all of its color when it becomes rancid. Some fats, however, contain traces of natural substances which tend to prevent rancidity, and such fats turn rancid very slowly and produce only a little acid.

Fats are complex substances made of glycerine combined with compounds known as fatty acids. Fats are broken down by the action of hot bases or acids or of superheated steam. If a base such as sodium hydroxide is used, the fatty acid will combine with it to form a salt. Soaps are fatty acid salts and are made by heating a fat with sodium hydroxide. If, for example, we treat a typical fat, such as the glyceryl stearate found in tallow, with sodium hydroxide, a reaction like this occurs:

glyceryl stearate (a fat) + sodium hydroxide → glycerine + sodium stearate (a soap)

If we treat soap solution with a strongly acid substance such as sodium bisulfate, the base will be taken away from the weak fatty acid, and the latter will remain in a free state.

EXPERIMENT 500—Making a Fatty Acid from Soap.

Dissolve a piece of soap the size of a pea in half a test tube of water by shaking the tube and heating. To the solution, add 3 measures of Sodium Bisulfate

(No. 7). The foam is instantly destroyed, and a thick white precipitate forms. This is the fatty acid which was in the soap. It can be separated by boiling for a minute or two, shaking gently at the same time to keep the liquid from boiling over. The fatty acid melts in the boiling liquid forming globules, which slowly collect at the surface Set the test tube aside to cool and the fatty acid will solidify into lumps. Remove these with a glass rod and use them in the next experiment. This fatty acid is mainly a mixture of stearic acid and palmitic acid.

EXPERIMENT 501—Making a Soap from Fatty Acid.

Place 2 measures of Sodium Carbonate (No. 4) and 2 measures of Calcium Oxide (No. 20) in a test tube and fill it ½ full of water. Boil this for a minute taking care not to let the hot liquid spurt out in your face, and then let the precipitate settle. The clear liquid is a solution of sodium hydroxide, a base.

Now put a small piece of the fatty acid which you prepared into another test tube and fill it about ¼ full with the sodium hydroxide solution. Boil it gently for a few seconds, and the lump of fatty acid will dissolve, forming soap. After cooling, close the mouth of the test tube with your thumb and shake vigorously. Notice the good lather which is formed. The reaction may be written as follows:

sodium hydroxide + stearic acid → sodium stearate (soap) + water

Most candles contain a little stearic acid to harden the parraffin wax so that it will not soften on warm days. You can try making soap out of a candle and if successful you will know that the candle contains stearic acid.

EXPERIMENT 502—Testing a Candle for Stearic Acid.

Put a few shavings from a candle into a test tube and add about ½ inch of sodium hydroxide solution prepared as in a previous experiment. Heat to boiling, and if foam or lather forms, the candle contains stearic acid. Pure paraffin gives no foam, only clear, oily globules.

EXPERIMENT 503—The Preparation of Acrolein.

The acrid odor of burning fat is due to a substance known as acrolein. It is also formed when a fat is broken down by heating with sodium bisulfate.

Put a small piece of fat such as butter in a test tube and add 2 measures of Sodium Bisulfate (No. 7). Heat carefully, pointing the mouth of the test tube away from you. After the butter has melted and begun to foam and boil, remove the test tube from the flame and smell cautiously the fumes which are given off. The penetrating odor is due to acrolein. Do not smell too much of it, for it is very irritating to the delicate tissues of the nose and eyes.

A few fats, in addition to glycerine and a fatty acid, contain also compounds of phosphorus and nitrogen. Egg yolk is rich in a fat of this type called lecithin. When heated with an alkali, its nitrogen is given off in the form of methyl amine, CH_3NH_2, a compound found also in decaying fish.

EXPERIMENT 504—The Formation of Methyl Amine from Lecithin.

Prepare some sodium hydroxide solution by boiling 2 measures of Calcium Oxide (No. 20) and 2 measures of Sodium Carbonate (No. 4) in ½ test tube of water

for a minute, being careful not to let the solution spurt out. Allow the white precipitate to settle.

Now put 2 or 3 drops of egg yolk into another test tube and fill it about ¼ full of the clear sodium hydroxide solution. Close the mouth of the test tube with your thumb and shake it well. Heat the mixture carefully, moving the test tube back and forth just above the flame, and shaking it gently, as suggested in the directions for heating in the front of this book. Boil the liquid carefully for a little while, and then remove the test tube from the flame and smell at its mouth. Notice the fishy odor of methyl amine.

BUTTER AND OLEOMARGARINE

When cream is agitated in a churn, the tiny droplets of fat in it collect into grains. These are separated from the liquid, that is, buttermilk, and a small amount of salt is thoroughly worked into it unless "sweet' (unsalted) butter is desired. The salt acts as a preservative and improves the keeping qualities of the butter. Usually dyes are also added to give a richer yellow color. Butter is perhaps the only food which may legally be artifically colored without that fact being stated on the label.

In addition to butter fat and salt, butter contains about 15% of water, distributed as tiny droplets throughout the fat. A small amount of milk protein, or curd, and milk sugar is also present.

Large amounts of renovated, or process, butter are made from rancid or spoiled butter by blowing air through it while hot to remove the rancid odor and then churning it with fresh milk.

Oleomargarine is a substitute for butter. It is made from hydrogenated cottonseed oil and other vegetable and animal fats by churning with a small amount of milk. The food value of oleomargarine is about the same as that of butter.

EXPERIMENT 505—Testing Oleomargarine and Processed Butter.

Place a small lump of butter in the spoon and heat over the alcohol lamp flame. If it is either oleomargarine or processed butter, it will sputter and crack like a green stick in a fire.

EXPERIMENT 506—Testing Fresh Butter.

Place a small lump of butter in the spoon and heat over a flame. If it is fresh butter, it will boil over, producing a large amount of foam.

EXPERIMENT 507—How Rancid Butter Can be "Renovated".

If you can obtain some rancid or spoiled butter, melt it in a small pan and pump air through it, using a bicycle pump, until there is no longer any odor noticeable. Now mix some sweet milk or cream with the butter and pour the whole mass into a screw top jar and shake the jar vigorously until the solid part which remains is very similar to fresh butter.

CARBOHYDRATES

Carbohydrates are substances made of various proportions of carbon combined in a certain way with hydrogen and oxygen. They always contain twice as much hydrogen as oxygen, the same proportions as in water. Starch, ordinary cane sugar, and cellulose, the woody fibrous material of plants, are familiar examples of carbohydrates.

The most important carbohydrates are those which are formed by green living plants and which are used as food by animals. The leaves of plants take up carbon dioxide from the air, and their roots absorb water from the soil. Sunlight shining upon the leaves causes these two compounds to combine, forming glucose or other kinds of sugar. This is transported to various parts of the plant and built into starch and cellulose. The process by which carbon dioxide and water, under the influence of sunlight, are converted into glucose, is one of the greatest mysteries of plant life, and no chemist has ever been able to make this reaction occur outside of a living

plant. It is one of the most important of all chemical reactions, since practically all living things depend upon it for food.

Cellulose is the most abundant carbohydrate. Wood is composed very largely of it, and the tough, fibrous parts of all plants owe their strength to it. In the form of grass and leaves, it is the principal food of horses, cows, sheep and many other animals. Although man himself is not able to digest cellulose, he uses it in a multitude of ways. In the form of wood, he uses it for buildings; in the form of cotton, linen and rayon fibers, for clothing; in the form of paper, it is one of the cornerstones of civilization.

STARCH

Starch is very similar to cellulose in chemical composition, differing only slightly in the way its elements are combined, but it is a much more easily digested carbohydrate, and hence is one of the most important foods of the human race. The greatest amount of starch is consumed as food just as it occurs in grains, roots, and fruits, but a considerable quantity is separated by mechanical processes and is used in laundry work, as a textile and paper sizing agent, as an adhesive, and for the preparation of dextrine and glucose.

Practically all plants contain starch, but potatoes, wheat, corn, and rice are the most convenient sources.

EXPERIMENT 508—Extracting Starch from Raw Potatoes.

Obtain a few slices of raw potato and mash them thoroughly in a dish or pan with a little water until the whole mass is reduced to a pulp. Add a little more water, stir well and then strain the mixture through a coarse cloth so as to remove the large fibers. Now place the part of the mixture which passed through the cloth in a tumbler, fill the tumbler about half full of water and stir the mass thoroughly. Stop stirring and allow the liquid to remain quiet for a few minutes. You will find that the white starch settles out first, forming a layer at the bottom, and on top of this will be some of the fibrous matter which was not taken out by straining. The water can be poured off the starch, and you can probably remove some of the remaining fibers of pulp with a spoon, leaving nearly pure starch.

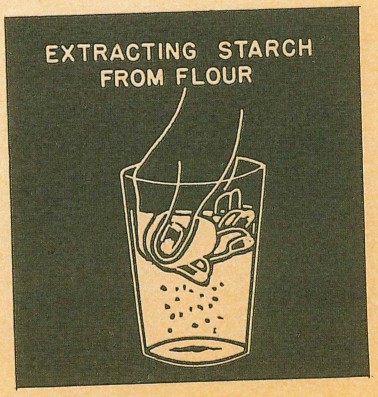
EXPERIMENT 509—Extracting Starch from Corn.

Crack open a few dry, uncooked kernels of corn and mash them in a dish or pan with a small amount of water until the mass is reduced to a paste. Strain this paste through a piece of coarse cloth to remove the husks, and then put the part which passed through the cloth into a tumbler and fill the tumbler half full of water. Stir this mixture for a few moments and then allow the solid matter to settle to the bottom of the tumbler.

The commercial extraction of starch is carried out along these same general lines. However, special apparatus is used which makes the operation easier and increases the yield of starch.

EXPERIMENT 510—How to Extract Starch from Flour.

Put 3 or 4 teaspoonfuls of wheat flour in a dish and add enough water to make a stiff dough ball, kneading it with your hands. Take a small piece of this dough ball about the size of a large marble, and hold it in a glass of water, at the same time working it between your fingers. You will find that the starch will separate from the flour and settle to the bottom of the tumbler, while what remains in your hand is the gluten. When no more starch seems to come from the dough or gluten, take another small piece in your fingers and work it in the water. In this way starch can be separated from flour.



EXPERIMENT 511—Dissolving Starch.

Starch is insoluble in cold water, but in boiling water it swells up, forming a paste which is called starch solution, although it is not a true solution.

The best way to dissolve starch is to mix a small quantity with cold water, then add boiling water to the mixture and heat it for a few moments longer. This makes a smooth starch paste, but if you were to try to dissolve starch by simply putting it in a little water and bringing it to boiling, it would form lumps and you would not get a satisfactory paste or solution.

EXPERIMENT 512—The Composition of Starch.

Put about 3 measures of ordinary cornstarch in a clean dry test tube and heat it gently over a flame. Notice that moisture is given off, as shown by the steam coming out of the tube. Presently the starch becomes brown, and if you continue to heat, it gradually turns black, and finally there is nothing left in the tube except a black mass of carbon. Thus we find that starch is composed of water and carbon. It is a compound very similar to sugar, except that the elements are present in a little different proportion, and its molecules are much larger.

Starch solution gives a characteristic blue color with iodine. This reaction is used to test for starch.

EXPERIMENT 513—Testing Flour for Starch.

Put a half measure of flour in a test tube half full of water and bring the solution to a boil. Now cool the solution and when it is perfectly cold, add a drop or two of tincture of iodine. A blue color will result, showing that starch was present in the flour.

Try this experiment with a little piece of bread and prove that it also contains starch.

EXPERIMENT 514—To Show that Potatoes Contain Starch.

Take a bit of potato, about the size of a pea, and boil it in a test tube one-third full of water for a minute or two, and then let the solution cool. The solution must be perfectly cold, otherwise the test for starch will not work. Now add 1 drop of tincture of iodine. A beautiful blue color will result, showing the presence of starch.

EXPERIMENT 515—Testing Rice for Starch.

Boil 2 or 3 grains of rice in a test tube half full of water for several minutes, and when the solution has become cold, add a drop of tincture of iodine. The blue color formed shows that starch is present in rice.

EXPERIMENT 516—Testing Corn Flakes for Starch.

Place 3 or 4 corn flakes in a test tube ¼ full of water and heat them to boiling over the flame of your alcohol lamp. Allow the solution to cool and then add 1 drop of tincture of iodine. The blue color which forms shows the presence of starch.

EXPERIMENT 517—Testing Corn for Starch.

Crush a grain of corn in a pan or dish. Place ½ measure of the crushed corn kernel in a test tube and add ½ inch of water. Heat the solution to boiling and then allow it to cool. Add 1 drop of tincture of iodine and again you will notice the blue color showing the presence of starch.

EXPERIMENT 518—Testing Oatmeal for Starch.

The oatmeal cereals are made of crushed, partially cooked seeds of oats. This material contains a great deal of starch. Place 3 or 4 flakes of oatmeal in a test tube ¼ full of water. Heat the water to boiling and allow to cool. Add 1 drop of tincture of iodine and you should get a deep blue color, proving the presence of starch.

EXPERIMENT 519—Testing Wheat Cereals for Starch.

If you can obtain any of the popular brands of wheat cereal in granular form, add about 1 measure to a test tube ¼ full of water. Heat to boiling over the flame of the alcohol lamp and then allow to cool. Upon adding 1 drop of tincture of iodine you will immediately notice the characteristic blue coloration which proves that the wheat cereal contains starch.

EXPERIMENT 520—Testing Unripe Fruit for Starch.

Obtain an unripe apple and cut it in halves. With a small brush or piece of absorbent cotton, paint one of the halves with tincture of iodine. Notice after a little while a blue color appears, proving that starch is present.

EXPERIMENT 521—Testing Ripened Fruit for Starch.

Unripe fruits contain starch, but as the fruit ripens the starch is converted into sugar. Repeat the preceding experiment, but this time use a fully ripened apple and observe that the tincture of iodine does not produce a blue color, which proves that ripened fruit does not contain starch.

Dextrine is made by heating starch. It is soluble in cold water and forms an excellent mucilage having a sweet taste. Everyone is familiar with the custom of starching clothes to make them stiff and glossy. When starched clothes are ironed, the heat from the iron converts some of the starch into dextrine, forming a glaze over the fabric.

EXPERIMENT 522—The Formation of Dextrine.

Put 4 or 5 measures of cornstarch in a clean dry test tube and heat gently over a flame. Water will first be given off, then the starch becomes brown, and an odor like that of bread crust is given off. Now stop heating the starch and examine the substance which remains in the tube. You have changed the starch into dextrine.

EXPERIMENT 523—How to Make Dextrine Paste.

Add a few drops of water to the dextrine formed in the above experiment and notice that the dextrine differs from starch in that it dissolves in cold water. Spread a little of this dextrine paste on a piece of paper and notice that it forms an excellent paste or mucilage. This is the purpose for which it is often used.

EXPERIMENT 524—The Iodine Test for Dextrine.

Shake a little of your dextrine paste in a test tube half full of water and add 1 drop of tincture of iodine. Instead of the characteristic blue color which starch gives with iodine, you will observe some shade of purple, red, or brown, depending upon how completely you converted the starch into dextrine.

SUGARS

When sugar is mentioned one naturally thinks of the sugar bowl. Ordinary cane or beet sugar is the most familiar member of the class of substances called sugars. The chemist calls it sucrose to distinguish it from other sugars.

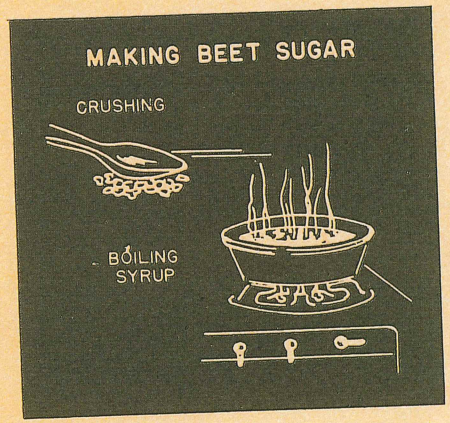
Sucrose is obtained principally from sugar cane and sugar beets. Sugar cane is extensively grown in the West Indies and most of our domestic sugar comes from these islands. In Europe, most of the sugar is made from sugar beets.

To obtain pure sugar from sugar cane, the stalks are first moistened and then passed between huge rollers to press out the sweet, sugary juices. Calcium or strontium hydroxide is added to remove certain impurities from the juices, and by evaporation, "raw" or brown sugar is crystallized out, and the dark thick liquid which remains is molasses.

To refine the brown sugar, it is moistened with sugar syrup, placed in a centrifuge and rotated at a high speed. This causes the impurities to be

thrown off, leaving the sugar crystals in a relatively pure state. These are further refined by dissolving in water and filtering through bone black or specially prepared charcoal. The filtrate is concentrated in vacuum evaporators, and crystallization is allowed to occur under conditions that control the size of the sugar crystals. After passing through driers it is ready for packaging.

Beet sugar is obtained by slicing sugar beets and allowing the sugar to diffuse out of them into hot water. The syrup thus obtained is refined in much the same way as that extracted from sugar cane, and the beet sugar finally obtained is exactly the same as cane sugar. At every stage in the manufacture of sugar, the chemist must make careful tests to regulate the process and to check the results.



EXPERIMENT 525 — Making Beet Sugar.

Obtain an uncooked sugar beet and cut it into thin slices. Put 3 or 4 of these slices into a saucer or bowl with a little water. Crush the slices with the back of a heavy spoon until you have reduced the mass to a pulp. Now taste the clear liquid which gathers on top of the beet pulp and see if you detect a sweet taste. Boil some of this clear liquid in a small pan until it becomes thick and syrupy. By putting this syrup in a cool place, sugar can be crystallized if the syrup is sufficiently thick and concentrated.

Like all carbohydrates, sugar is composed of carbon, hydrogen, and oxygen. When it is heated, the hydrogen and oxygen are driven off in the form of water, and finally only carbon is left. If only part of the hydrogen and oxygen are driven off, a substance known as caramel is left.

EXPERIMENT 526—How to Make Caramel.

Put about 10 measures of sugar in a test tube and heat it very gently over a flame for a few minutes. Notice that the sugar melts and that steam is given off. In a short time the mass becomes brown and you can smell caramel. Remove the test tube from the flame and examine the caramel. This substance is used as a flavoring and coloring material for candies and various sweets. It is also used for coloring vinegar and beverages.

EXPERIMENT 527—Combustibility of Sugar.

Sugar burns just like wood. In fact its composition is very similar to that of wood. Great care must be taken in large sugar refineries to guard against fire. Place a piece of lump sugar on a tin pan and apply heat beneath it. Notice how easily the sugar catches fire and how readily it burns.

EXPERIMENT 528—Luminescence of Sugar.

Rub 2 pieces of ordinary lump sugar together in the dark and notice the faint flashes of light which are given off. Of course the room in which you try this experiment must be absolutely dark or you will not be able to see the flashes. Luminescence produced by rubbing is called triboluminescence.

Dozens of other sugars in addition to sucrose are known to the chemist. Some of them occur in nature; others are synthetic products made in the

laboratory. Some of the more common sugars are glucose (dextrose, or grape sugar), fructose (levulose, or fruit sugar), lactose (milk sugar), maltose (malt sugar), and xylose (wood sugar).

Of these, glucose is the most abundant in nature and the most important commercially. It is made on a large scale by heating starch with dilute hydrochloric acid under pressure. It is the principal sugar present in corn syrup, and it is used for sweetening all sorts of commercial food products—candy, jelly, preserved fruit, baked goods, beverages, and so forth.

EXPERIMENT 529—How to Make Glucose.

Put 10 measures of cornstarch and 3 measures of Sodium Bisulfate (No. 7) in a test tube. Fill the tube half full of water and gradually heat the liquid until it boils. Continue to boil for 5 or 10 minutes. Now pour a few drops of this solution into another test tube, add an equal amount of water and then wait until this solution has become thoroughly cool. You can run cold water over the bottom of the test tube if you wish to hasten the cooling. When the solution has become cold, add 1 drop of tincture of iodine. If a blue color is formed, it shows that some of the starch still remains and it will be necessary to continue to boil the solution in the other test tube a little longer. If no blue color results, all of the starch has been converted into glucose. Pour the glucose solution into a shallow saucer and allow it to cool. After it has stood for a little while, solid glucose will begin to crystallize if the air is sufficiently dry.

EXPERIMENT 530—Conversion of Starch to Glucose by Digestion.

The starch in starchy foods is converted into sugar during the process of digestion in our small intestines. This reaction can be illustrated as follows:

Prepare a small amount of starch paste by mixing a few measures of starch with a little water in a test tube. Then add half a test tube of boiling water to the starch cream and boil the liquid for a few minutes. Now collect about ¼ test tube full of saliva. If necessary, chew a piece of paraffin or rubber for a few minutes to promote the flow of saliva. Add about ¼ test tube of the starch paste to the saliva and keep the mixture warm for a few minutes. The bottom of the test tube may be immersed in a pan of warm water. Do not heat it over a direct flame as too high a temperature would destroy the complex substances in the saliva which bring about the digestion of the starch.

After the mixture of starch paste in saliva has stood for a while, pour a few drops into another test tube. Cool this sample and add 1 drop of tincture of iodine. On the first test, you will probably get a blue color showing that the saliva has not completely digested the starch. Allow the test tube containing the mixture to remain in the warm water for a while longer and then test a few drops again. After a while you will find that you are getting a reddish or brown color, due to the presence of dextrine, and finally you will find that you do not get any color at all, showing that the starch has been converted into glucose by digestion with the saliva.

Pectin, which is used to make jelly set, belongs to a class of substances called glucosides. These are complex relatives of the sugars, which occur in the skin of most fruits and vegetables. Crab apple skins are especially rich in pectin, hence crab apples are much in demand for jelly-making.

EXPERIMENT 531—Making Pectin from Crab Apples.

Crush 4 or 5 crab apples in a cloth by pounding them with a hammer. Be sure to catch all of the juice in a pan, then strain the juice twice through a fine cloth to remove the pomace or pulp. Fill a test tube about ¼ full of the juice and add ½ test tube of alcohol. You can use rubbing alcohol, radiator alcohol, or ordinary denatured alcohol. At once you will notice a cloudy gelatinous precipitate beginning to form. This is crude pectin which must be purified for use in the manufacture of jellies and preserves.

EXPERIMENT 532—Testing Currants for Pectin.

Red currants are rich in pectin. Obtain about ½ cupful of red currants and crush them with the back of a spoon. Strain the mashed currants through a piece of fine cloth and catch the juice in a tumbler. To ⅛ test tube of this juice add about ⅛ test tube of water and shake the test tube vigorously. Next add about ½ test tube of alcohol and notice the precipitate of pectin which forms.

PROTEINS

The food constituents known as proteins are complex compounds of carbon, oxygen, hydrogen, nitrogen, and sulfur, with smaller amounts of phosphorous and occasionally other elements. All living cells, including those of our own bodies, contain proteins, and it appears that without them, life is impossible.

The proteins in the food which we eat are broken down by the gastric juice of the stomach and the digestive fluids of the small intestines into simpler compounds of the same elements. These are readily absorbed through the walls of the intestines and are carried by the blood to all parts of the body where they are again built up into proteins.

Typical foods composed largely of proteins are the lean portions of meat, the white of eggs, the gluten of flour and the curds of milk.

By heating with lime, the nitrogen which all proteins contain is driven off, largely in the form of ammonia.

EXPERIMENT 533—Testing Meats for Proteins.

Put a small bit of lean meat in a test tube, add 1 measure of Calcium Oxide (No. 20) and 2 or 3 drops of water. Heat this mixture and test for ammonia. Logwood paper is very good for making the ammonia test. A piece of moistened logwood paper will turn blue in the presence of ammonia.

Of course you can also test for the ammonia by smelling or by holding a moistened piece of red Litmus Paper (No. 18) at the mouth of the test tube. Often, however, the odor of ammonia is masked by that of other substances.

EXPERIMENT 534—Testing Milk for Proteins.

Place 4 or 5 drops of milk in a test tube and add 1 measure of Calcium Oxide (No. 20). Heat the mixture gently. Smell at the mouth of the test tube and you will immediately detect the odor of ammonia. You can also test for this gas with a piece of red Litmus Paper (No. 18). The paper will be turned blue by the ammonia.

EXPERIMENT 535—Testing Potato Skins for Proteins.

Potato skins are of course vegetable matter, but they contain a small amount of proteins. Put a small piece of potato skin in a test tube together with 1 measure of lime and 2 or 3 drops of water. Heat this mixture and see if you can show the presence of ammonia.

EXPERIMENT 536—Proving that Eggs Contain Proteins.

Place in a test tube a small amount of some white of an egg, together with 1 measure of Calcium Oxide (No. 20) and 2 or 3 drops of water. Heat this mixture gently and smell at the mouth of the tube for ammonia. Also test for this gas with a piece of moistened red Litmus Paper (No. 18).

EXPERIMENT 537—Proving that Hair Is a Protein.

Put a small wad of hair in a test tube, add a few drops of water and 1 measure of Calcium Oxide (No. 20), and heat gently. Occasionally remove the test tube from the flame and test with red litmus paper at its mouth for the ammonia which should be given off. We do not ordinarily think of hair as a protein, and of course it is not digestible, but chemically, it is very similar to the familiar edible protein materials.

EXPERIMENT 538—Proving that Horn Is a Protein.

Repeat the preceding experiment using instead of hair a few scraps of animal hoofs or horns, or even a few clippings from your own finger nails. Hair, nails and horn are composed largely of a tough, insoluble protein known as keratin. Wool, being hair, is also made of keratin.

EXPERIMENT 539—Proteins Contain Hydrogen and Oxygen as Water.

Dry a small amount of white of egg or a small piece of lean meat in the air without heating it. When perfectly dry, put the substance in a test tube and heat over a flame. Notice that moisture condenses on the upper part of the test tube, showing that the proteins contained hydrogen and oxygen in the form of water.

EXPERIMENT 540—Proteins Contain Carbon.

Heat a portion of protein, either lean meat or white of egg, in a test tube until water vapor is no longer driven off. Notice that the mass remaining is black and consists chiefly of carbon.

EXPERIMENT 541—Proteins Contain Sulfur.

Mix some white of egg with enough Calcium Oxide (No. 20) to make a dough so that you can roll the mass into balls with your finger. Put this mixture in a dry test tube and heat. The mass turns brown and gives off an unpleasant odor. Let the test tube cool and then fill it ¼ full of water. Shake thoroughly and pour off the brown liquid into another test tube. Add 1 or 2 measures of Sodium Bisulfate (No. 7) and heat to boiling. Remove the test tube from the flame and smell at its mouth for hydrogen sulfide. Test also with a moistened piece of Sulfide Test Paper (No. 31). The brown stain of lead sulfide shows that the egg white contained sulfur.

EXPERIMENT 542—Proving that Hair Contains Sulfur.

Put a small quantity of hair in a test tube ¼ full of water. Add 1 measure of Calcium Oxide (No. 20) and 1 measure of Sodium Carbonate (No. 4) and boil the solution for several minutes. Cool the solution. Now put about 2 measures of Sodium Bisulfate (No. 7) into the tube. It may take a little more than 2 measures as the sodium bisulfate should be added until no more effervescence occurs. Heat the solution again and test the steam which is given off with a small strip of Sulfide Test Paper (No. 31). The black stain which appears on the paper is caused by hydrogen sulfide gas; which proves that there must have been sulfur in the hair.

Proteins are divided into a number of classes. The albumins are an important class of proteins which is distinguished from the other classes by the ease with which heat precipitates them from their water solutions. Egg white contains a considerable amount of albumin.

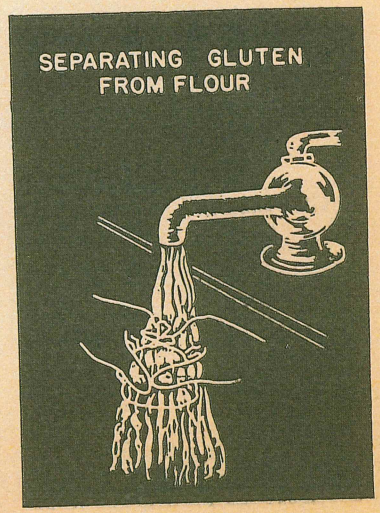
EXPERIMENT 543—Effect of Heat on Albumin.

Place 2 drops of egg white in a test tube and fill the tube half full of water. Shake thoroughly and then heat gently over a flame. Before the water boils, you will see a white cloud form in the solution. This is what happens when an egg is cooked. The albumin is coagulated to solid particles.

When the starch is washed out of wheat flour, an elastic substance known as gluten remains. It is a mixture of two proteins, glutenin and gliadin.

EXPERIMENT 544—Separating Gluten from Wheat Flour.

Gluten gives wheat flour the property of forming a light, fluffy bread. It may be separated from the flour in the following manner: Mix 2 teaspoonfuls of flour with enough water to form a stiff dough. Wrap this dough in a piece of cloth and knead the mass in your hands under a stream of water until the water which drains off is free from starch. The elastic material which remains in the cloth is gluten, and its quality determines to a large degree the quality of the flour. Good gluten is yellow and elastic, while gluten from old flour is gray and brittle.



EXPERIMENT 545—The Preparation of Casein.

One of the principal proteins found in milk is casein. Obtain a small quantity of skim milk and dilute it with 3 or 4 parts of water. Now dissolve 5 measures of Sodium Bisulfate (No. 7) in a test tube half full of water and add this solution to some of the skim

milk contained in another test tube. Shake the milk every time a little of the sodium bisulfate solution is added.

You will notice that a precipitate is formed in the milk and upon allowing it to stand for a while the precipitate becomes heavier. It consists principally of casein. Casein is used as an adhesive and it is sometimes mixed with formaldehyde, pressed into a hard stone-like mass and used as a substitute for ivory.

Casein is the base for cheese. When the milk ferments or sours, lactic acid is formed. This acid precipitates the casein just as the sodium bisulfate did in this experiment.

MILK

Milk is used in large amounts directly as a beverage, indirectly in the preparation of many other foods, and industrially as a source of casein. Powdered milk is made commercially by evaporating whole milk to dryness. Great quantities of powdered milk are used in the manufacture of bread, ice cream, and candies.

EXPERIMENT 546—Making Powdered Milk.

Obtain about ½ cup full of whole milk and place it in the inside container of a double boiler. Fill the outer vessel of the double boiler about ⅓ full of water and place over the burner of a gas or electric range. Fit the inside vessel into the outer one. After the water in the boiler begins to boil, stir the milk occasionally during the next half or three-quarters of an hour. At the end of this time you will notice that the milk has completely dried. The dried milk can be removed and ground to a fine powder. Powdered milk is made commercially by spraying the liquid milk into a current of hot air so that evaporation is almost instantaneous and the powdered milk is collected at once. In making this material you have evaporated nearly all of the water (about 88%) from the milk and thus formed "powdered milk."

Caution: When performing this experiment, be sure that you do not allow the boiler to become dry. Keep sufficient water in it at all times during the experiment.

EXPERIMENT 547—Making Condensed Milk.

Put about a cup of milk into a clean pan and heat it carefully on the stove. After the milk has boiled for a few minutes you will find that it is becoming thicker and it must be carefully stirred so that it will not scorch and stick to the pan. When the milk has been thickened or condensed to the desired point it can be sweetened by adding a little sugar. Commercially, condensed milk is made by low temperature evaporation in vacuum pans.

EXPERIMENT 548—Detecting Foreign Matter in Milk.

Filter a quantity of well-mixed milk and examine carefully the sediment which remains on the filter paper. If possible, use a magnifying glass and you will be able to see distinctly any dirt or foreign matter which was in the milk. A wad of cotton in the funnel is often used instead of filter paper.

TESTING FOODS

The Federal government maintains chemical and bacteriological laboratories in most of the large cities throughout the nation for the purpose of testing foodstuffs. Thousands of tests are conducted each year in cooperation with local health boards to see that consumers obtain food conforming to proper standards.

Adulteration of foods was at one time very general, and is still frequently practiced, although the pure food laws have largely done away with the more harmful forms. There are several reasons why food is adulterated.

One of the chief forms of food adulteration is artificial coloring matter. Sometimes poor food products are colored to look like first-class goods. In other cases artificial products are colored to imitate the natural products. For example, oleomargarine is often colored to represent real butter. Occasionally foods are colored simply to make them look attractive, for example, candies and fancy pastries.

Another reason for adulterating food products is to preserve them. Of course, many foods do not keep sweet and fresh very long, so various chemicals are added to keep them from spoiling.

Still another reason for adulterating foods is to increase the bulk or weight. For example, chicory or ground peas are sometimes mixed with coffee so that if you buy a pound of coffee which has chicory mixed with it, you get only about three-fourths of a pound of pure coffee.

EXPERIMENT 549—Testing Rye Flour for Wheat Flour.

Obtain 2 small pieces of window glass or other flat pieces of glass. Put a small quantity of water on one of the pieces of glass, enough to cover a space about ½ inch square, and put 2 measures of flour to be tested in the water. Spread this mixture around and press the other piece of glass down upon it. If wheat flour is present, white spots will be seen, and if the glasses are slid back and forth upon each other, the spots will pull out into threads.

EXPERIMENT 550—Testing Bread for Alum.

Prepare a solution of logwood by heating 1 measure of Logwood (No. 29) in a test tube ½ full of water. Filter this solution and then moisten a slice of bread with it. If the bread is free from alum, it will have a reddish brown color, but if alum is present in the bread it will become a lavender or blue color in 2 or 3 hours. The best way to carry this out is to pour a little of the logwood solution on a slice of bread, as the color can then be observed very plainly. Alum was formerly added to flour sometimes to conceal poor quality, but if alum is now found to be present in bread, it is almost certainly from an alum baking powder and not from adulteration of the flour.

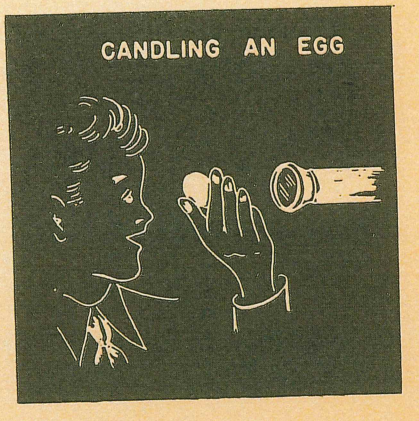
EXPERIMENT 551—Testing Ground Meat for Starch.

Sausage, deviled meat, and similar products are sometimes mixed with cracker and bread crumbs to add bulk and weight. The best way to detect this adulteration is to test for starch. Boil a small quantity of the meat in a test tube half full of water, and when the liquid has become cool add 1 drop of tincture of iodine. A blue coloration shows the presence of starch. A very small amount of starch may be present in meat products in the form of spices, but if a strong coloration is obtained, it is safe to assume that the adulteration is in the form of bread or cracker crumbs.

EXPERIMENT 552—How to "Candle" Eggs.

Candling is one of the methods most frequently used to determine the freshness of an egg. Hold an egg between the eye and an artificial light in a darkened room. A fresh egg should appear unclouded, homogeneous, and almost translucent. If dark spots are found, it is stale. A rotten egg appears dark colored.

Against the larger end of a fresh egg, between the shell and the lining membrane, a small air cell should be distinctly visible. In an egg which is not perfectly fresh, this space is filled with egg substance, unless the egg has been stored with the large end up.



EXPERIMENT 553—Testing an Egg in Salt Solution.

Since the density of an egg decreases by the evaporation of moisture through the porous shell, its freshness may be approximately estimated by placing it in brine. Prepare the salt solution by dissolving 12 spoonfuls of salt to 1 pint of water. Immerse the egg in the solution. A perfectly fresh egg will sink; if

several days old, it will swim just immersed in the liquid; if stale, it will float on the surface.

EXPERIMENT 554—Another Test to Determine the Freshness of Eggs.

Shake an egg, holding it near the ear. The contents of a fresh egg should not move. If the slightest movement can be detected, it is somewhat stale; if it rattles, the egg is spoiled. Open the egg and observe the odor and taste. If there is a tendency for the white and yolk to run together, the egg is not fresh, or the hen has been improperly fed.

EXPERIMENT 555—Testing Baking Powder for Alum.

Some baking powders contain alum as the acid ingredient. Prepare a solution of logwood by putting 1 measure of Logwood (No. 29) in a test tube ¾ full of water. Heat this solution to boiling and then filter it. Add 4 measures of baking powder to a test tube full of water, and when the effervescence has stopped, add 3 or 4 drops of vinegar. Now add 2 or 3 drops of logwood solution, and if alum is present, a bluish red color will appear.

EXPERIMENT 556—Testing Baking Powder for Starch.

Starch or flour is frequently added to baking powder. This substance cannot be considered as an adulterant as it prevents the ingredients of the baking powder from reacting before the powder is used.

Put 1 measure of baking powder in a test tube ½ full of water and heat the solution to boiling. Allow it to cool and then test with 1 drop of tincture of iodine. A blue color shows that the baking powder contained starch.

EXPERIMENT 557—Testing Canned Foods for Iron.

Canned fruit is sometimes acid. This may be the natural fruit acid, or some acid may have been added to preserve the fruit. When the fruits are canned, it sometimes happens that there is a defect in the tin coating on the inside of the can, in which case the acid may attack the iron and dissolve a little of it. This is not harmful, but may make the food taste.

To test for iron in canned fruit, pour a little of the juice into a test tube and add a few drops of solution prepared by dissolving one measure of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. A blue coloration indicates that the fruit juice contains some iron.

EXPERIMENT 558—How to Test Honey for Gelatin.

Dissolve a small portion of honey in a test tube ¼ full of water and heat the solution until it becomes thick and syrupy. If the honey was adulterated with gum or gelatin, the mass will gelatinize on cooling. Prepare a solution of ½ measure of Tannic Acid (No. 15) in a test tube ½ full of water and add a few drops of this to the solution of honey. If gelatin is present in the honey, a precipitate will be formed with the tannic acid. Honey containing gelatin, when heated with Calcium Oxide (No. 20), gives off ammonia gas.

EXPERIMENT 559—Testing Ground Coffee.

Ground coffee is sometimes adulterated with ground chicory, beans or peas. Sometimes such a mixture can be seen by means of a magnifying or reading glass.

Place a half teaspoonful of ground coffee in a glass of water and stir for a moment. All of the impurities and a little of the coffee will sink to the bottom, but most of the coffee will float. If there is chicory in the mixture, the particles which sink will leave a trail of dark brown colored liquid behind them.

Coffee itself does not contain starch, while the substances used to adulterate it do. Hence, the test for starch may be used to show the presence of adulterants. Enough water should be added to the solution of coffee so that the brown color will not disguise the blue color of the starch iodine reaction.

EXPERIMENT 560—Testing Lemon Exract for Lemon Grass Oil.

Lemon extract is sometimes adulterated by using oil of lemon grass as a substitute for genuine lemon oil. Lemon oil does not mix with water. Put about ½ inch of lemon extract in a test tube half full of water. Shake the mixture and if it becomes clear, it shows that no lemon oil is present and the extract is adulterated. If the mixture is cloudy or turbid, however, it shows that the extract contains the genuine oil of lemon.

EXPERIMENT 561—Testing Mustard for Starch.

Mustard is sometimes adulterated with mustard hulls, rice, or wheat flour, and when any light colored flour is used, a yellow coloring matter is generally added

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EXPERIMENT 561—Testing Mustard for Starch.

Mustard is sometimes adulterated with mustard hulls, rice, or wheat flour, and when any light colored flour is used, a yellow coloring matter is generally added

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* * *

PRICE LIST OF CHEMICALS AND APPARATUS

Additional supplies of any chemicals or extra apparatus can be obtained by sending directly to The Porter Chemical Co., Hagerstown, Maryland.

Send checks, money orders, or cash with your order. Please include 10c for postage on orders of 25c or less, include 20c for postage on orders between 25c and $1.00, and include 25c for postage on orders between $1.00 and $2.00. We pay the postage within the United States on all orders for $2.00 or over. Foreign Postage and Duties, including Canadian Duty, are extra in all cases.

Refill chemicals are packed in the same style containers as in the original outfit. In ordering refills state the number of your outfit and the edition of your CHEMCRAFT book. Coins sent by mail should be very carefully wrapped, and unless letter is registered are at sender's risk. U. S. Stamps will be accepted for amounts up to $1.00. Canadian or foreign stamps or coins will not be accepted.

SUPPLIES FOR CHEMCRAFT CHEMISTRY OUTFITS

THIS IS NOT THE LIST OF THE CONTENTS OF ANY OUTFIT

1—Sulfur	$ .10
2—Sodium Borate	.10
3—Trisodium Phosphate	.10
4—Sodium Carbonate	.10
5—Sodium Bisulfite	.10
6—Sodium Ferrocyanide	.15
7—Sodium Bisulfate	.10
8—Dye Wafer	.10
9—Ammonium Chloride	.10
10—Strontium Chloride	.15
11—Zinc Metal	.10
12—Aluminum Sulfate	.10
13—Nickel Ammonium Sulfate	.15
14—Tartaric Acid	.20
15—Tannic Acid	.20
16—Sodium Thiosulfate	.10
17—Ferrous Ammonium Sulfate	.15
18—Litmus Paper	.10
19—Sodium Silicate Solution	.15
20—Calcium Oxide	.10
21—Ferric Ammonium Sulfate	.15
22—Phenolphthalein Solution	.15
23—Sodium Iodide Solution	.15
24—Manganese Sulfate	.10
25—Sodium Bicarbonate	.10
26—Sodium Salicylate	.15
27—Cobalt Chloride Solution	.15
28—Chameleon Paper	.10
29—Logwood	.10
30—Mixed Dyes	.10
31—Sulfide Test Paper	.10
32—Congo Red Paper	.10
33—Glycerine	.15
34—Blue Print Paper	.10
35—Tumeric Paper	.10
36—Powdered Iron Metal	.10
37—Tin Metal	.20

38—Calcium Hypochlorite	.10
39—Azurite	.15
40—Ammonium Molybdate	.25
41—Starch	.10
42—Phenolphthalein Paper	.10
43—Iron Sulfide	.10
44—Iron Pyrites	.10
45—Chrome Alum	.15
46—Ammonium Hydroxide	.15
47—Cochineal	.15
48—Magnesium Sulfate	.10
49—Calcium Sulfate	.10
50—Calcium Carbonate (Precipitated Chalk)	.10
51—Copper Metal	.10
52—Gum Arabic	.15
53—Carbon Tetrachloride	.10
54—Boric Acid	.10
55—Powdered Charcoal	.15
56—Charcoal Lump	.15
57—Calcium Chloride	.10
58—Flame Test Wire (Nichrome)	.10
59—Blank Test Paper	.10
60—Filter Paper	.10
61—Stearic Acid	.15
62—Benzoic Acid	.15
63—Graphite	.10
64—Sodium Ammonium Hydrogen Phosphate	.15
65—Aluminum Metal	.15
66—Bismuth Subnitrate Solution	.15
67—Potassium Carbonate	.10
68—Dilute Hydrochloric Acid (5%)	.15
70—Dilute Acetic Acid	.15
75—Dimethylglyoxime Solution	.15
Eight-hole test tube rack without drying pegs	.30
Spoon	.05
Measure	.05
Test Tube Brush	.10
Test Tube Holder	.10
Watch Glass	.20
Tripod	.50
Wire Gauze (Asbestos Center) 4 x 4 inches	.20
Mortar and Pestle (1 ounce)	.50
Flask (Florence) 60 ml.	.35
Flask (Erlenmeyer) 125 ml.	.55
Flask (Florence) 250 ml.	.55
Beaker (Resistant Glass) 50 ml.	.35
Beaker (Resistant Glass) 100 ml.	.45
Funnel	.35
Gas Generating Apparatus Complete	.50
Gas Generating Apparatus Without Funnel	.25
Gas Delivery Tube and Stopper	.10
Wash Bottle Assembly Complete (250 ml. flask, No. 4 rubber stopper, 3 pieces of bent glass tubing)	.80
Stirring Rod	.05
Glass Tube	.05
Assorted Glass Tubing (7 pieces)	.25
Test Tubes (4 x ½ inch size)	5c ea., 40c per doz.
Test Tubes (5 x ⅝ inch size)	10c each, 3 for 25c
Rubber Tubing ($\frac{1}{16}$ inch Inside Diameter) per foot	.07
Two-hole Rubber Stopper, No. 4	.07
Chemcraft Alcohol Lamp Blow Torch	.50
Alcohol Lamp (2 oz.) Complete with Wick and Metal Cap	.50
Alcohol Lamp (1 oz.)	.25
Alcohol Lamp Wick	.05
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